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EXPERIMENTS

in Laboratory Chemistry

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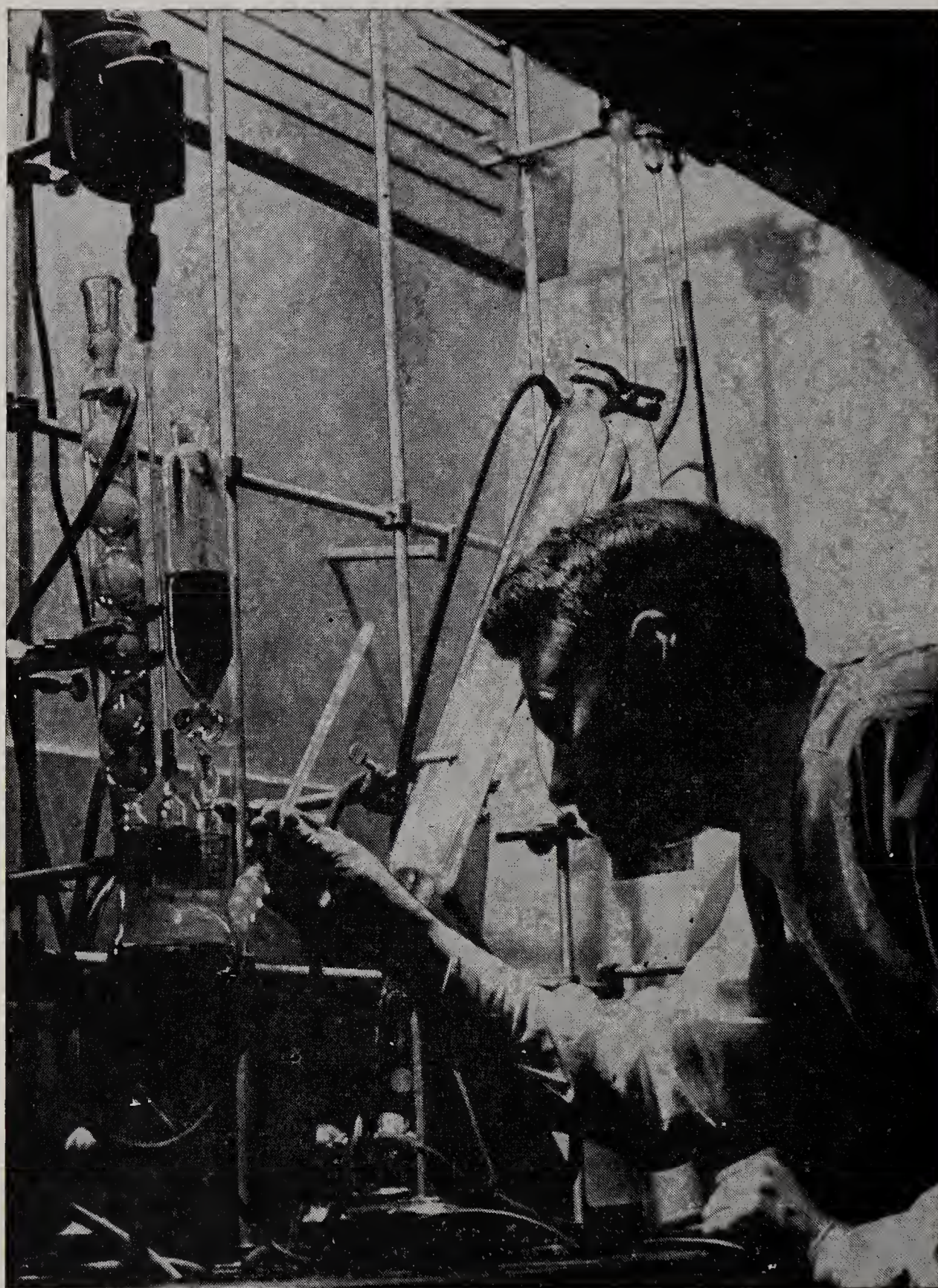
COLLEGE SAINT JEAN

EDMONTON, ALBERTA

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Experiments in Laboratory Chemistry

M. Blackburn
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Observe carefully and measure accurately.

EXPERIMENTS

in Laboratory Chemistry

To Accompany

CHEMISTRY FOR SECONDARY SCHOOLS

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Preface

"Experiments in Laboratory Chemistry" is designed for the beginner in chemistry and is especially intended for use with the text "Chemistry for Secondary Schools" by the same authors. The student is advised to use the text only to clarify and co-ordinate the experimental work of the laboratory.

This book contains more experiments than can be carried out in a year's work. It is suggested that, from the material provided, the teacher will select those experiments that are most suitable for his classes and that can be performed with the apparatus available. It is better to have a smaller number of experiments performed carefully, thoroughly understood, and well written, than to rush through many experiments carelessly. Experiments should not be omitted if their omission in any way interferes with the continuity of the course.

For most of the experiments, only simple inexpensive apparatus is needed. Experiments designated "*(Demonstration)*" are either too costly, too difficult, or too dangerous to be performed by beginners, and should be carried out on the lecture table, either by the teacher or by assisting pupils under close supervision.

A mere mechanical performance of an experiment has little value. Consequently, the student should become familiar with the background of each experiment by carefully reading the exercise beforehand. Where an experiment is involved, or lengthy, the reading assignment should be given on the day previous to the laboratory period.

Complete and precise directions are given which are so clearly stated and illustrated, that the teacher is not burdened with the necessity of dictating details of manipulation and procedure.

Following each procedure is a series of questions (in italics) designed to direct the attention of the student to the important features to be observed. Questions are also asked at the end of some experiments and units, where it is necessary to gather together the observations of several procedures or experiments in order to make generalizations.

Chemistry is essentially an experimental science. The laboratory work gives reality to the subject by presenting concrete sensory experiences which aid the student in grasping important ideas, and makes the fundamental principles a part of his own experiences. The laboratory is the best place in which to develop a truly scientific attitude.

A. G. C.

J. H. C.

A. H. L.

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General Laboratory Instructions

The following suggestions are intended for the guidance of the student:

1. Familiarize yourself with the name and general purpose of each piece of apparatus in your locker.

2. Neatness and cleanliness are essential if you are to do accurate work. Keep your locker orderly and your table-top and apparatus clean. Experiments performed with dirty equipment often produce misleading results.

3. Before you leave the laboratory be sure that the gas and water taps are turned off, that all your equipment is clean and in its proper place, and that the desk-top is sponged clean, and dried. A single drop of acid may cause a severe burn or damage the clothing of a student in the class that follows.

4. Solids such as matches, litmus papers, broken glass, metals, and unconsumed insoluble chemicals should never be placed in the sink, where they may be the cause of clogged plumbing. Put these articles in the waste jars provided for that purpose. *Soluble* substances and liquids (with the exception of mercury) may be poured into the sink, but should be flushed out with plenty of running water.

5. **Follow directions.** Before you begin an experiment, the directions should be read through carefully until the method of procedure is clearly understood. Then follow the directions closely.

6. **Mistakes are dangerous.** Always be certain that you have the correct material (or materials) for the experiment. Take time to read the labels on reagent bottles carefully. Using the wrong material may cause painful burns. Warning will be given in such cases and there is little danger if you follow the directions closely.

7. **Quantities of materials.** You will save time and achieve better results if you use the quantities of materials called for in the directions. Chemicals are expensive and the results of the experiments are often obscured by using *too much*. Actually weigh or measure these quantities until you have learned to closely estimate a gram or a millilitre.

8. **First Aid.** Familiarize yourself with the First Aid Instructions to be found in the appendix at the end of this book.

Any injury or accident, no matter how trivial it may seem, should be reported immediately to the teacher in charge.

9. **Laboratory techniques.** In a later section, certain simple operations and manipulations are discussed, which you will be required to carry out in laboratory practice. Possibly the best method is to learn each particular operation at the time it is first encountered in the experimental procedures. This method should be used thereafter on all occasions. TECH. (*Technique*) has been frequently inserted at various places in the text material to caution you to turn to this section and read the particular technique to which your attention has been directed.

Laboratory Notebooks

The chief benefits that a student should derive from an elementary course in chemistry are training in careful observation, practice in the proper manipulation of apparatus, and in accurate description and interpretation of experimental results. The keeping of a good, well-organized notebook, therefore, is an essential part of a student's training.

It is not necessary that all experimental records conform to a rigid pattern; rather, after some initial guidance has been given, the student should be encouraged to organize his own notebook. Certainly, the study of chemistry should not be sacrificed to meticulous conformity to mechanical notebook procedure.

Although loose-leaf notebooks that accommodate the various subjects taken by the student are commonly used, many teachers prefer a separate chemistry notebook. A single hard-covered book (approximately 8" x 11") containing approximately 100 pages of unruled paper is selected by many teachers. This may be fitted with a liner ruled by the student to meet his individual writing style.

The following are a few suggestions to the student:

1. Do not duplicate the printed instructions in the book. Instead, state in your own words, (a) what was actually done, (b) what was actually observed, and (c) the conclusions that were drawn from the facts revealed by the experiment. Be careful to draw only those conclusions that are warranted by the facts.

2. Arrange your records in an orderly manner and write in ink. Record the number and title of each experiment and the date on which the experiment is performed. Each record should be so clear

and complete, that any person reading it would have no difficulty in understanding the experiment.

3. A well-labelled cross-sectional diagram will make your written description of the experiment easier to understand. The diagrams should be large enough for clarity and understanding. Good judgment should be exercised in the time spent on making diagrams.

4. Answer each *italicized question* with a definite statement, in such a manner that no difficulty will be encountered in recalling the question.

5. Where the experimental work does not provide sufficient information to answer a question adequately, a page reference to the text has been given, indicating where the necessary information may be obtained. The text, however, should not be consulted until the experiment has been performed.

Percentage of Error

The difference between your experimental result and the accepted value which has been established by trained scientists making a large number of trials with high-grade equipment is called the *actual error*. A much more significant expression of the degree of accuracy is to find what fraction (expressed as a percentage) this actual error is of the accepted value. For example, the gram-equivalent weight of magnesium is 12.12 grams and your experimental result is 11.86 grams. Your actual error is 0.26 grams, which seems at first inspection to be a sizable error. The *percentage of error* is $\frac{0.26}{12.12} \times 100$, or 2.14%.

(To find the percentage of error, divide the actual error by the accepted value, and multiply the quotient by 100.)

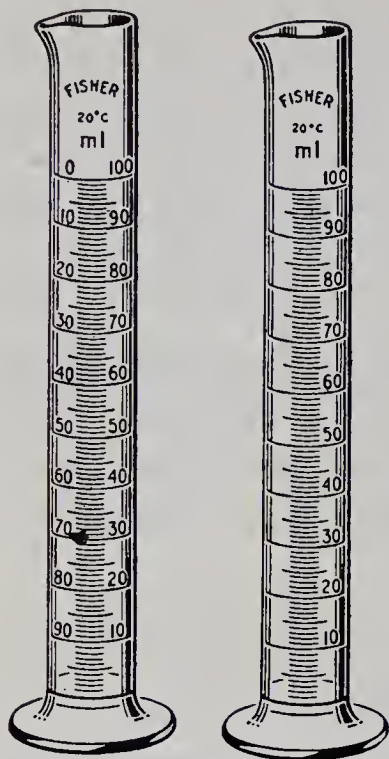
It is impossible to establish a uniform rule for the degree of accuracy to be expected in the various quantitative experiments covered in this manual.

Laboratory Techniques

A. Working with liquids

1. **Measuring volumes.** (a) *Graduated cylinders.* The volumes of liquids are commonly measured in cylindrical glass vessels with

graduated scales on their sides. They are often called "chemical graduates" and are obtainable in various capacities. The scale



Fisher Scientific Company

Fig. 1. Graduated cylinders.

etched in the glass shows the volume in millilitres (cubic centimetres), or in some cases in fluid ounces. Some have a single scale in which the divisions are numbered from the bottom upwards, enabling the observer to measure the volume of liquids *poured into* the graduate. Others have a double scale, where the divisions in this added scale are numbered from the top downward, thus making it possible to measure the volume of liquids *poured from* the graduate. (Fig. 1).

You should familiarize yourself with the scale markings on the cylinders available in your laboratory, and their values. You should also note that the top surface of an enclosed liquid is curved, not flat. This curved surface is called the *meniscus*. With liquids such as water, that wet the glass walls, a concave meniscus is observed. In reading the volume of such liquids, care must be taken to have the graduate on a fixed, level surface and your eye on a horizontal line passing through the lowest point on this curved surface. Figure 2 A illustrates the correct method of reading a graduate.

(b) *Burettes*. These are calibrated tubes of various capacities, with the zero mark found near the top (Fig. 2B). A glass stopcock near the bottom controls the flow of liquid from the tube. They may be filled by placing a funnel in the top of the tube and adding the liquid until it rises above the zero mark. The stopcock (tap) is then opened and sufficient liquid is allowed to escape until all the air in the tip has been driven out and the upper level of the liquid is *at* or *below* the zero mark. Any desired volume may then be released by careful manipulation of the tap. The initial and final position of the meniscus is read in the manner described earlier.

(c) *Pipettes*. The pipette is a glass tube tapered to a point at one end and commonly having a bulb-like enlargement near its centre. It is calibrated to deliver a specified volume of liquid at the temperature indicated. The pipette is filled by suction until the liquid is slightly higher than the graduation mark found on the stem. Then

the upper end of the tube is quickly covered with the finger. By a slight release of the pressure exerted by the finger, the liquid is allowed to descend slowly until the bottom of the meniscus curve is coincident with the mark on the stem, as shown in Fig. 2C.

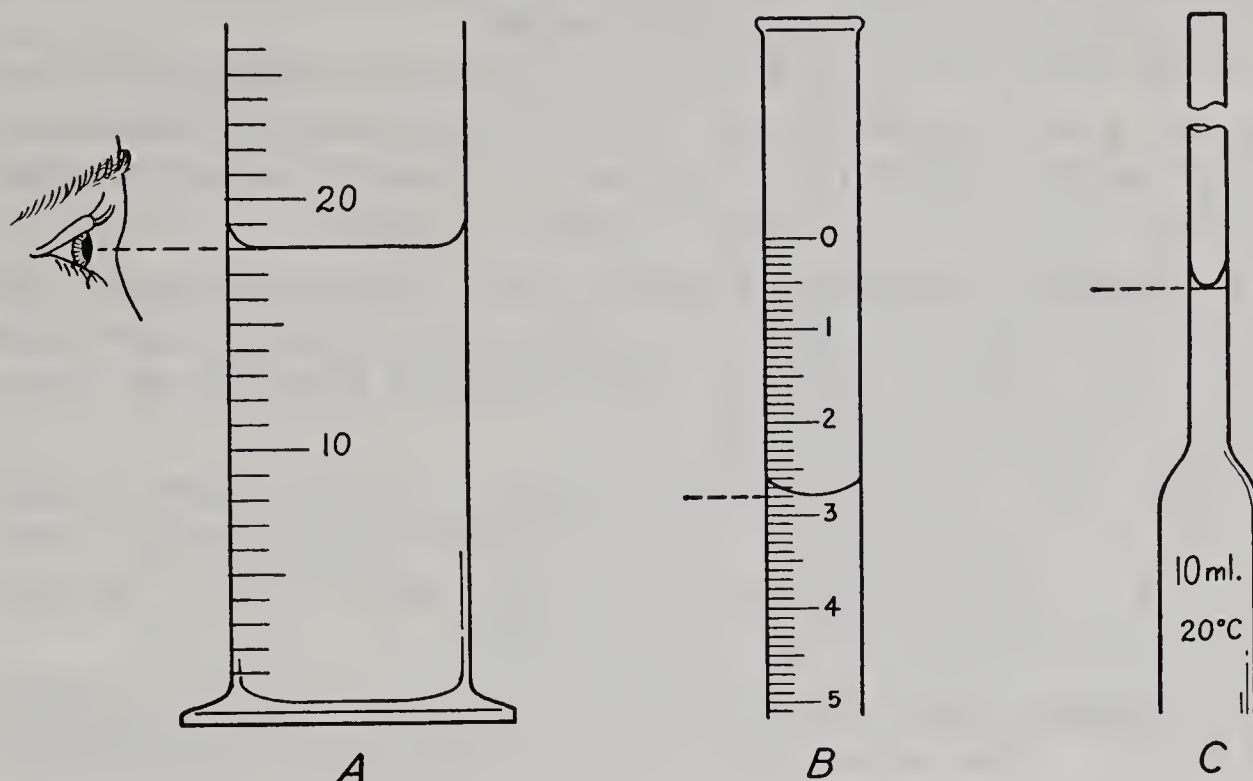


Fig. 2. Measuring volumes of liquids. A, using a graduated cylinder; B, using a burette; C, using a pipette.

When the pipette is allowed to deliver its specified volume of liquid, the last few drops which remain in the pointed end are released by drawing the point along the wet wall of the vessel into which the liquid is being delivered.

2. Using reagent bottles. When pouring a reagent from a glass-stoppered bottle, *the stopper must never be placed on the desk*. It may become contaminated, and when returned to the bottle, make

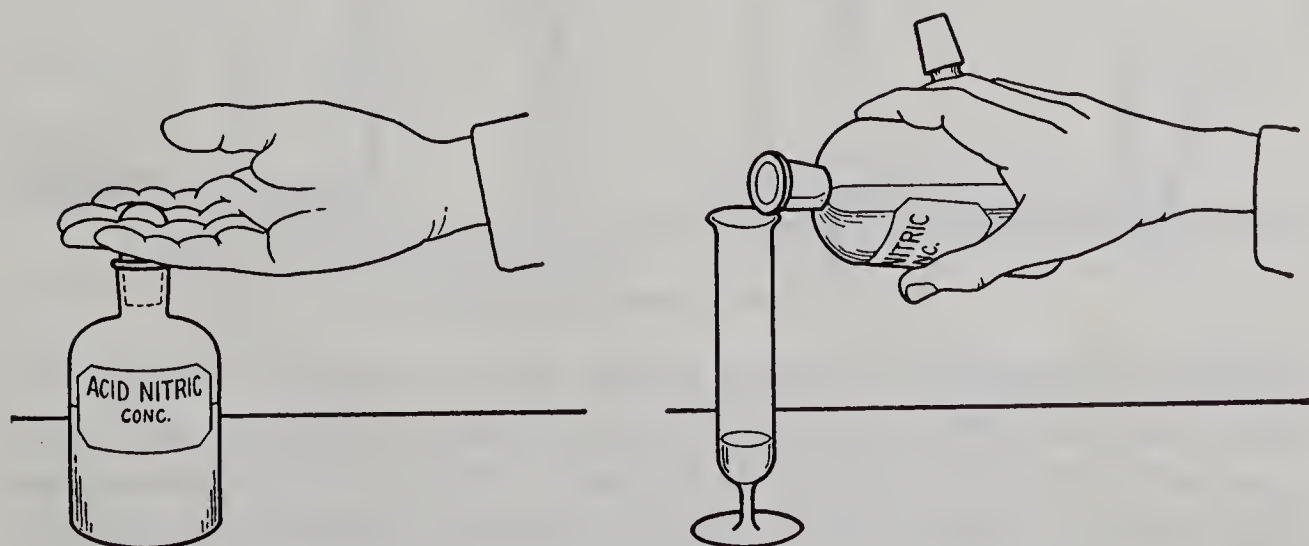


Fig. 3. Pouring a reagent from its container.

the contents unfit for further use. In the case of strong acids and bases, it may leave a drop of the reagent on the desk top which can mar the finish, damage someone's clothing, or cause painful burns. Always hold the stopper between the fingers, as indicated in Fig. 3, leaving your hand free to grasp the bottle.

When pouring from the bottle, the neck of the bottle should *touch* the rim of the receiving vessel. This prevents the liquid from trickling down the side of the bottle and is especially important when handling acids, bases, or other corrosive liquids.

3. Diluting concentrated sulphuric acid. *Never pour water into concentrated sulphuric acid.* To dilute concentrated sulphuric acid with water, *the acid should be slowly added to the water*, in a thin stream, with constant stirring.

When water and concentrated sulphuric acid are mixed, a great quantity of heat is evolved. If the water is erroneously poured into the acid, a rapid formation of steam may occur, causing the contents of the container to be violently ejected.

4. Filtering. Insoluble material may be separated from a liquid by filtering. The porous filter paper is prepared for the funnel by

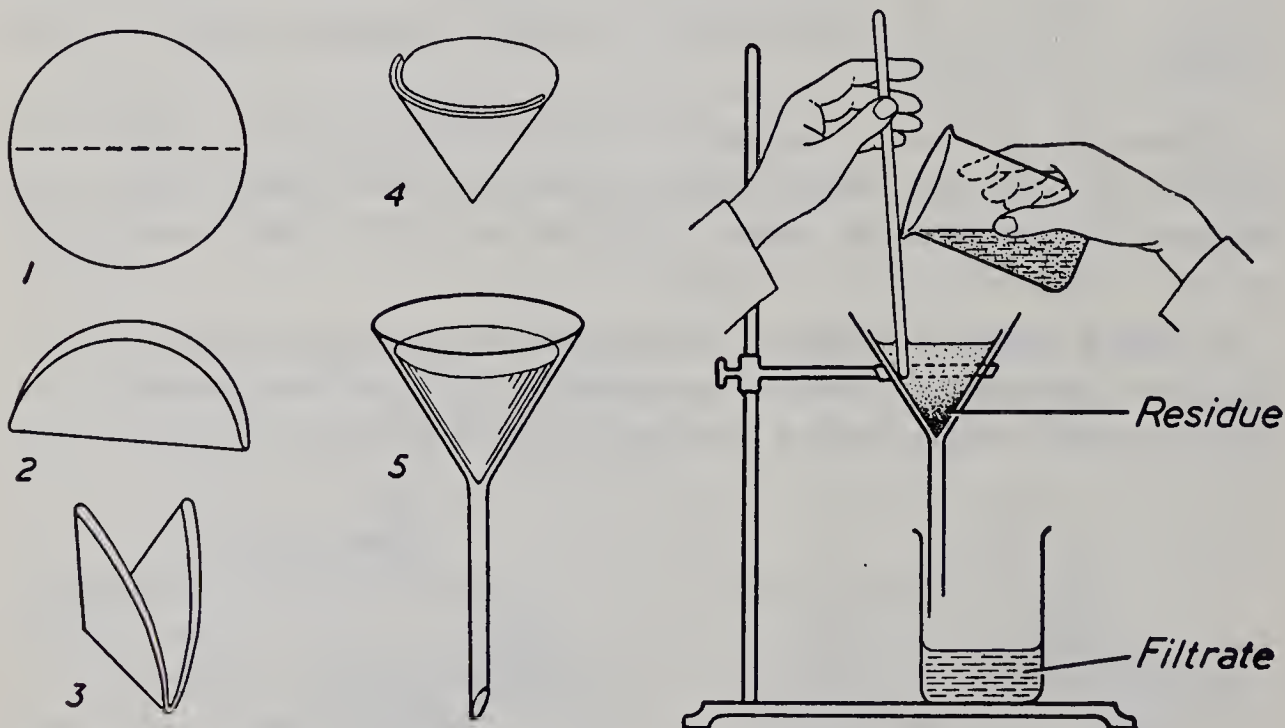


Fig. 4. The process of filtration.

folding it successively into the shapes shown in Fig. 4. When the paper has been folded along its diameter and then again at a right angle to the first fold, one of the outside folds is opened so that the paper assumes a cone-like shape with three thicknesses of paper on one side and one thickness on the other. This is placed in the

funnel and sufficient water added to make it adhere to the glass (with no air spaces), so that it will filter rapidly. This treatment also discourages the absorption of large quantities of the solution being filtered. After the moist paper has been pressed down with the fingers, any excess water should be discarded before using the filter. The paper should not extend above the edges of the funnel, but its apex should always project slightly into the stem.

Support the funnel in a ring attached to a retort stand and place an empty beaker or test-tube under the funnel to catch the *filtrate*. The liquid to be filtered should slowly be poured down a glass rod into the funnel, as shown in the accompanying diagram. The lower end of the rod should touch the filter paper inside the funnel, so that the liquid will run down the side and thereby avoid breaking the apex of the filter paper. The acute edge of the funnel stem should touch the side of the collecting vessel. This permits the filtrate to run down the side of the receiving vessel and prevents spattering.

After the liquid has drained from the solid material (*residue*) retained by the paper, soluble impurities are often removed by *washing* the residue with several 5 ml. portions of distilled water, and allowing each portion to run through before adding the next. (TECH. 9)

5. Evaporating over a water bath. If the evaporating dish containing the solution to be evaporated is placed on a wire gauze supported on a retort stand, and slowly heated over a low flame, "spattering" commonly occurs as the condition of dryness is approached. This causes the loss of some of the final product.

Usually it is advisable to support the evaporating dish over a beaker of boiling water

as illustrated in Fig. 5. The dish is heated by the steam produced. Care must be taken to replace the water in the beaker from time to time, since allowing it to boil dry is likely to crack the glass.

In some non-quantitative experiments where only a small volume of a solution is to be evaporated to dryness, the "water bath" method

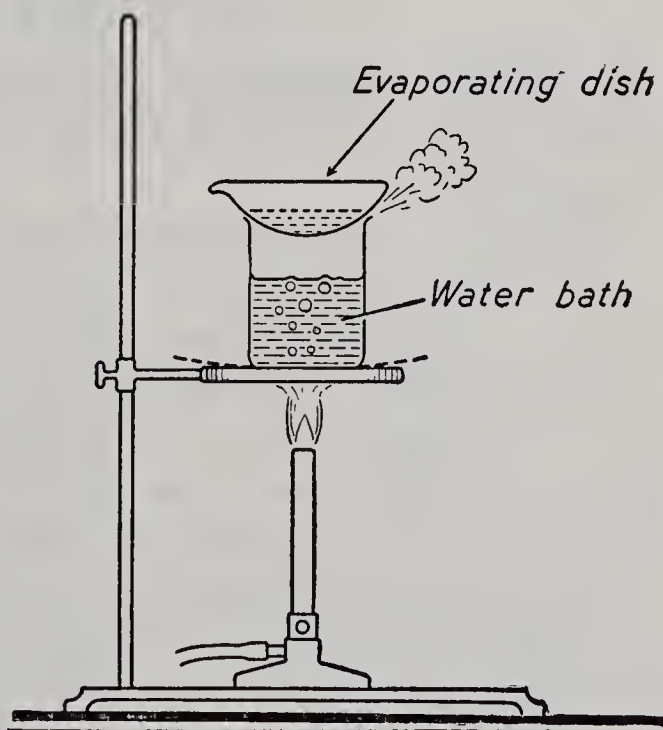


Fig. 5. Using a water bath.

is not used. Instead, the evaporating dish containing the solution is supported on wire gauze and is heated strongly at first until the volume of liquid is considerably reduced. Then, the final evaporation to dryness is accomplished with a much lower flame, since small pockets of steam cause the liquid to "spatter" as the solution becomes more concentrated.

When only a few drops of a solution are to be evaporated, a watch-glass is usually employed. This should be encircled with the thumb and first finger and held over a low flame. Unless it is held thus, it may be brought too close to the flame and the watch-glass may crack.

6. Testing with litmus. Except where a different method is specifically stated, litmus tests on *liquids* should be conducted by placing the paper (or papers) on a clean glass plate and transferring to the paper with a clean glass rod, a drop or two of the liquid to be tested.

When *gases* are to be tested, the litmus paper (or papers) should first be moistened with water.

Never handle litmus papers with wet or dirty fingers, nor permit them to lie on the desk top.

B. Working with solids

7. Heating glassware. (a) Every piece of apparatus made of glass or porcelain should be thoroughly dry on the outside before being heated.

(b) The temperature should be raised gradually. Sudden heating or cooling may cause the apparatus to crack. (c) Beakers, flasks, evaporating dishes, glass retorts, etc., should not be heated in direct contact with the flame. Support them on a wire gauze with an asbestos centre. A porcelain crucible and its lid that have been warmed gradually in the initial heating and allowed to cool slowly, may be supported on a clay triangle and heated directly.

(d) Never clamp flasks or test-tubes more firmly than is necessary to hold them in position. The expansion of the glass on being heated may break it if it is clamped too tightly.

(e) Test-tubes may be heated with a direct flame. A strip of paper several times folded, may be substituted for a wire holder if these are not available.

In heating a test-tube containing a liquid, it should always be held in an *inclined position*. The temperature should be raised *gradually* and *uniformly* to avoid the sudden ejection of the contents because of chemical action, or the formation of steam. This **uniform**

heating may be accomplished by moving the tube back and forth through the flame, and gently agitating the contents of the tube. Always direct the mouth of the test-tube away from yourself or anyone near while the heating is in progress.

8. Putting a solid into a test-tube. Cut a piece of stiff, glazed paper an inch longer and slightly wider than the diameter of the test-tube used. Fold it lengthwise to form a V-shaped trough. Place the solid near one end of the trough. Hold the test-tube in a horizontal position and slide the trough inside the tube, as shown in

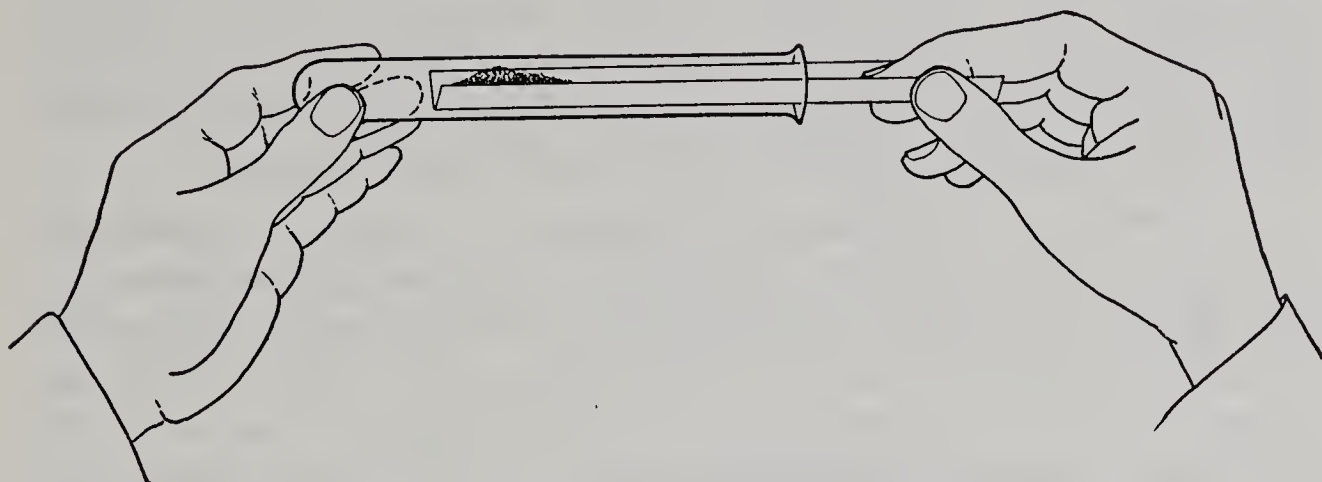


Fig. 6. Introducing a solid into a test-tube.

Fig. 6. By inclining the tube and gently tapping the paper, the material will be deposited in the bottom of the test-tube, leaving the sides clean.

9. Washing a precipitate. When an insoluble substance is filtered out of a solution, the pores of the filter paper retain a portion of the solution itself. Some of the solution also adheres to the particles of retained solid. To eliminate all traces of the soluble material, the residue is washed by adding distilled water. The wash bottle shown in Fig. 7 is commonly used as a container for the distilled water. On blowing into the mouth-piece, a fine stream is forced out of the jet. Sometimes the water is poured directly from the mouth-piece in 5 to 10 ml. portions, and each portion is allowed to run through the filter paper before adding the next. This treatment is continued until all the soluble material is washed through. The insoluble material can then be dried and weighed.

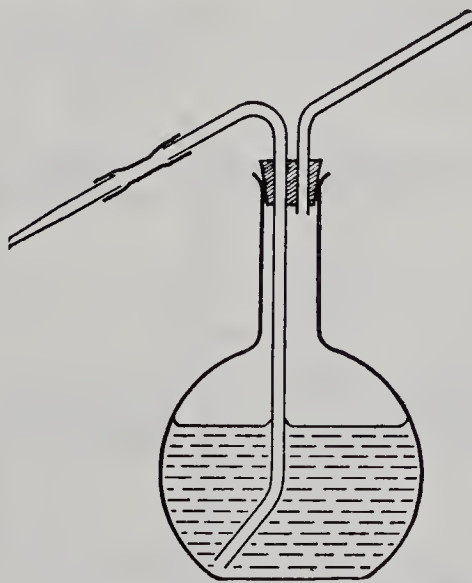


Fig. 7. A wash bottle.

10. Handling yellow (white) phosphorus. Yellow phosphorus is stored under water because it has a low ignition temperature and takes fire on exposure to air. It causes severe burns if allowed to come in contact with the skin. Consequently *it must always be cut under water and must be handled with forceps; never with the fingers.*

If cut fragments are made available at the teacher's table, the student must keep the phosphorus under water while carrying it to his desk.

Any traces of unburned phosphorus remaining on a deflagrating spoon must be burned before putting the spoon away. This is accomplished by heating the spoon to redness for two minutes. Phosphorus deposits on flasks and test-tubes may be safely removed by treatment with an acidified solution of bleaching powder.

11. Inserting glass tubing into stoppers. It must be remembered that glass is fragile and that carelessness in this simple operation can easily result in a badly lacerated hand. Always lubricate the end of the tubing and the channel in the stopper with water. Hold the stopper between the thumb and forefinger of one hand and with the other hand grasp the tubing *with a towel* at a point approximately one inch from the entering end. Use a gradual rotary motion with the towel-wrapped hand and shift the grip as often as necessary.

C. Working with gases

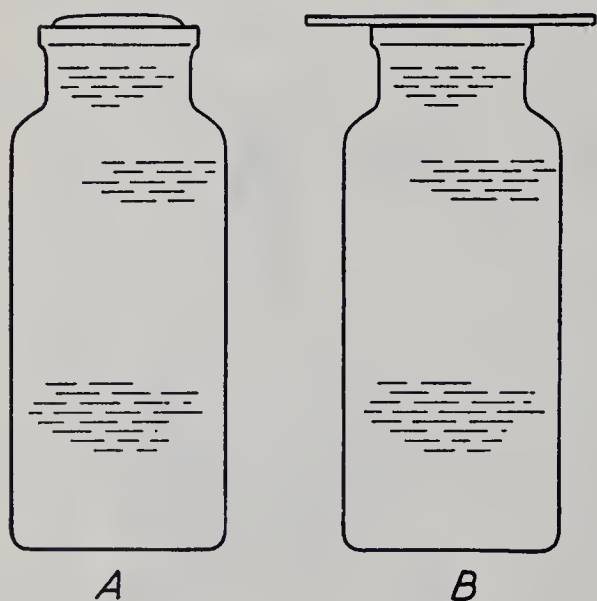


Fig. 8. The proper method of filling a gas bottle with water in collecting a gas by the downward displacement of water.

12. Collecting gases by displacement of water. When gases are to be collected over water, the gas bottle should be filled with water until the rounded upper surface of the water extends above the mouth of the bottle, as shown in Fig. 8A. A glass plate should then be slid over the mouth of the bottle and held tightly in place while the bottle is being inverted over water in a pneumatic trough, or the sink. The glass plate is removed as soon as the mouth of the bottle is below the surface of the water in the trough.

To collect the gas, the bottle should be tipped enough to permit placing the end of the delivery tube under its mouth. When rubber

tubing is used as the delivery tube, care must be taken not to pinch the tubing with the tipped bottle. The use of a "beehive", support as shown in Fig. 9A, eliminates the need for tipping the bottle, since it has an opening in the top over which the water-filled gas jar is rested. An arched opening in the base permits the insertion of the tubing through which the gas passes. The pneumatic trough provided in your laboratory may be equipped with a removable, perforated, sliding shelf, which functions in the same manner as a beehive support.

In experiments where the flask or test-tube in which the gas is being generated is heated, care must be taken never to allow the generator to cool while the free end of the delivery tube is still under water. As the generator cools, an internal decrease in pressure occurs. The cold water forced back into the generator by the greater atmospheric pressure may crack the generator.

13. Collecting gases by displacement of air. If the gas being collected is appreciably soluble in water, collection by displacement

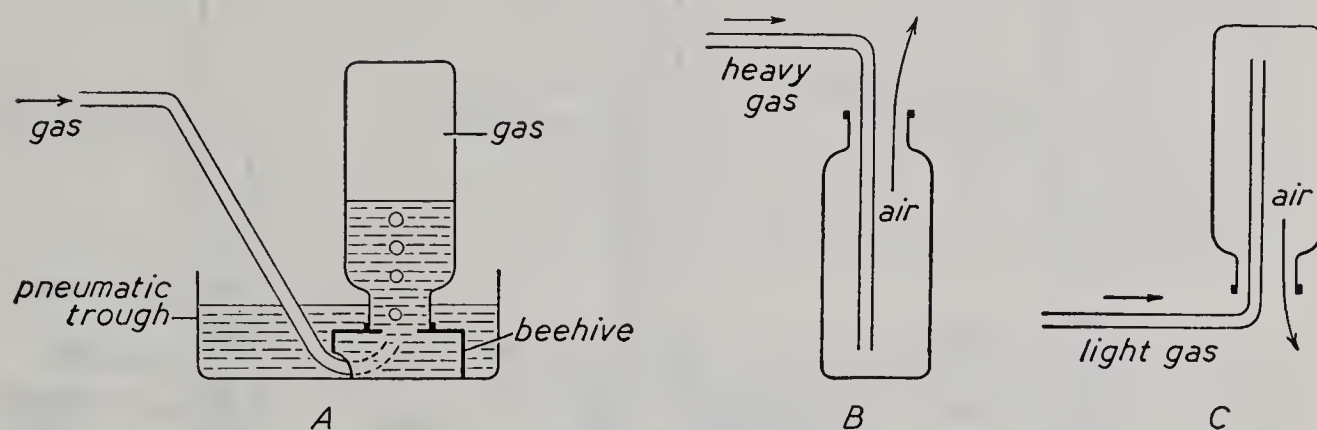


Fig. 9. Three methods of collecting a gas. A, downward displacement of water. B, upward displacement of air. C, downward displacement of air.

of water has to be discarded. If the gas is heavier than air, it is collected by the *upward* displacement of air, as shown in Fig. 9B. If the gas is lighter than air, it is collected by the *downward* displacement of air, Fig. 9C. The delivery tube should be extended to the base of the collecting bottle in both cases.

14. Smelling a gas. Use care in smelling gases. Keeping your mouth closed, wave the gas toward your nose with your hand, and sniff gently.

15. Making splint tests. Wooden splints of cedar or pine are frequently used to determine whether a particular gas burns or supports combustion. They should be 8 to 10 inches long, and about $\frac{1}{4}$ -inch in diameter. Some experiments demand the use of a *glowing* splint;

in others a *blazing* or flaming splint is required. A glowing splint is prepared by igniting the splint in a Bunsen flame and allowing it to blaze until there is a spark of glowing charcoal at the tip when the flame is shaken out.

The Bunsen Burner

The Bunsen burner is almost indispensable as a source of heat and is used in all laboratories where some type of gaseous fuel is available. Various models are available, but all are so constructed that a mixture of air and the gaseous fuel can be ignited at the top of the vertical barrel. The nature of the flame and the temperature produced are controlled by the relative volumes of gas and air admitted.

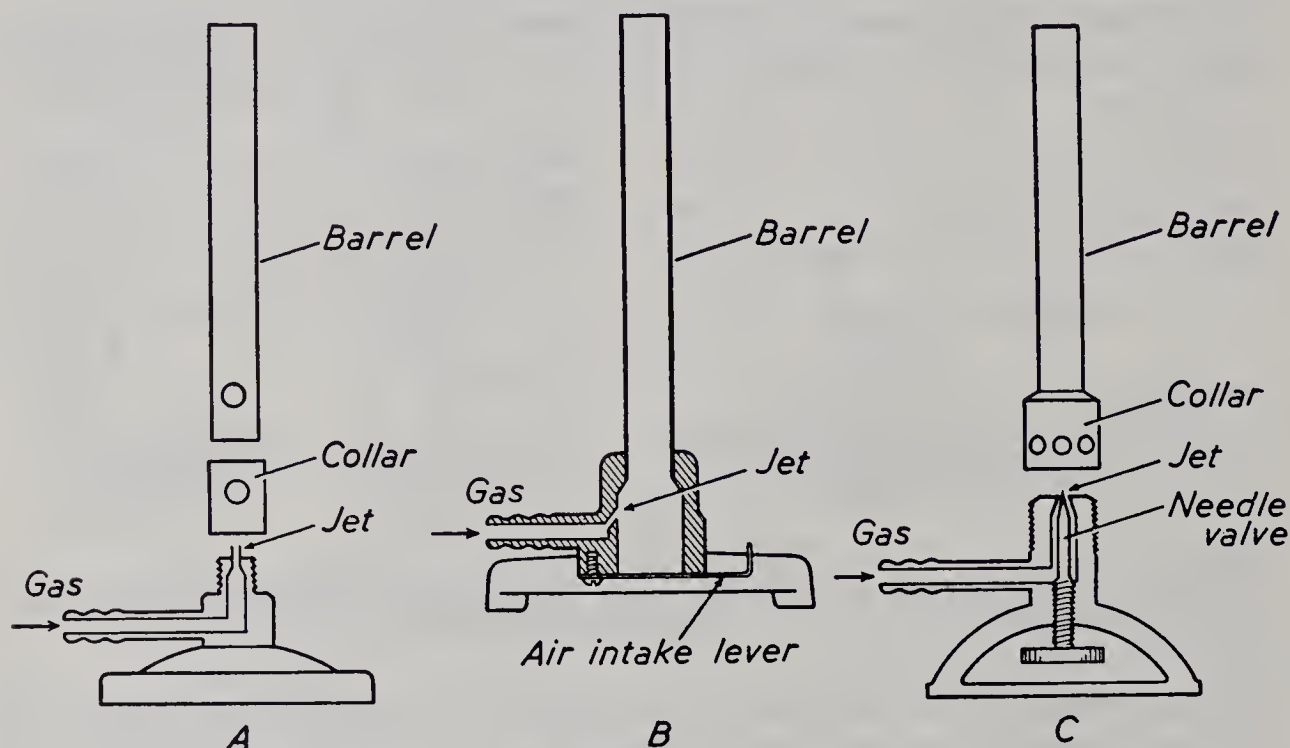


Fig. 10. Various types of Bunsen burners.

Model (A) in Fig. 10 represents one of the commonest types. The barrel may be unscrewed from the base, thus revealing the small opening in the jet through which the gas enters the mixing chamber. The collar is a perforated ring which fits over the base of the barrel in such a manner that the openings in the barrel may be kept open or closed by rotating the collar. The burner is connected to a gas outlet by means of thick-walled, tightly-fitted rubber tubing.

Model (B) has an H-shaped base and the orifice which admits the gas is displaced to the side wall of the stack, the latter being firmly

seated in the base casting. An air-intake adjustment of special design is operated by a lever on the underside of the base.

In model (C) the air inlets are located in the base of the barrel and the gas flow is controlled by a knurled adjustment screw which regulates a needle valve.

In some models, the gas chamber, orifice, air jets, and hose connection are all contained in the base casting and a slight rotation of the barrel adjusts the air supply. Possibly your laboratory may have *Meker* or *Fisher* type burners in which the top of the barrel is fitted with a metallic grid. The latter prevents back-flashing and produces a flame that is quite different from the Bunsen flame. Instead of the long narrow Bunsen flame with a hot tip and a cold central cone, a wide flame of great intensity that is uniformly hot throughout, is produced.

Examine the burner provided in your laboratory. Note the features of its construction and make a labelled diagram of the burner.

Lighting the burner. Attach the rubber tubing of the burner to a gas outlet and control the air supply by partially closing the air inlet. Light a match, turn on the gas almost full, and after a few seconds light the burner by bringing the match across the top of the barrel. Do not hold the lighted match over the burner before turning on the gas, since the match will probably be extinguished before the gas ignites.

High and low flames. Since the burner operates most efficiently when the mixture of air and gas is such that the flame is neither coloured yellow nor noisy, you should now adjust the air supply until all traces of yellow disappear, but not far enough to make the flame noisy. Establish a quiet, non-luminous (blue) flame displaying a distinct inner cone. This is the type of flame you will employ in your experiments unless directed otherwise.

The inner cone is hollow, the space in its interior being occupied by an unburned mixture of air and gas. The object being heated should be held near the top of the upper cone, slightly below the tip. If a *high* flame is specified, the stopcock on the gas pipe should be regulated until a flame not more than 4 or 5 inches high is obtained. When a *low* flame is specified, the gas supply is reduced.

Learn how to create a low flame and a high flame.

Draw the upper portion of the barrel of your Bunsen burner and show a properly adjusted high flame with its cones. Indicate the region where an object should be placed if it is to be heated to best advantage.

A luminous flame. When the air supply is inadequate, the flame produced is sooty, luminous, and not very hot. Cut off the air supply on your burner by closing the air inlet and note the yellow flame produced. Using a pair of tongs, hold a porcelain crucible cover or a glass plate in this flame. The black deposit is carbon resulting from the incomplete burning of the gas, and the yellow colour of the flame is due to the heating to incandescence of unburned carbon particles. You should never use this type of flame unless specifically directed to do so.

Back-flashing. If too much air is admitted, the flame sometimes “*strikes back*” and burns at the jet near the base of the burner. When this happens, turn off the gas supply at once and let the burner cool. Then *decrease* the amount of air admitted to the burner and relight. It can also be controlled by adjusting the flow of gas entering the burner. This back-flashing is dangerous because of the poisonous gases released in the laboratory. It can also be the cause of severely burned fingers.

Working with Glass

Cutting. Determine the length needed, lay the tubing on the desk, and at the place where it is to be cut make a short but deep

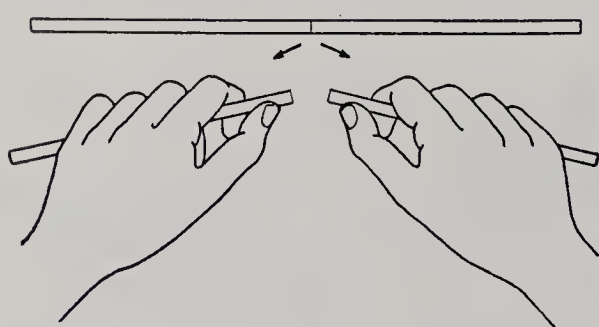


Fig. 11. Cutting and breaking glass tubing.

transverse scratch with a forward stroke of a triangular file. Then grasp the tubing firmly with a hand on either side of the scratch, holding the thumb tips together behind the mark. Now press outward with the thumbs, pulling gently inward at the same time with the hands, and the tube should break at the point where

the scratch was made (Fig. 11). If the tubing does not yield readily to gentle pressure, a deeper scratch must be made.

The sharp edges should be rounded off (fire-polished) by slow rotation in the Bunsen flame at a downward angle of 45° until the glass begins to melt slightly. Relatively smooth edges may be rubbed with emery cloth or a file instead of being fire-polished.

Bending. Glass tubing should always be bent in a *flattened* flame. A Bunsen flame can be flattened by the use of a “wing top” attachment which slips over the top of the burner tube as shown in Fig. 12.

For ordinary bends the flattened flame should be slightly yellow and about two inches in width.

A *right-angle* bend is made as follows. Determine the point at which the tubing is to be bent. Grasp the tubing in both hands and hold slightly above the flattened flame so that about 5 or 6 centimetres of the tube is heated where the bend is to be located. Slowly *rotate* the tubing between the thumbs and forefingers to insure uniform heating and gradually lower it into the position shown in

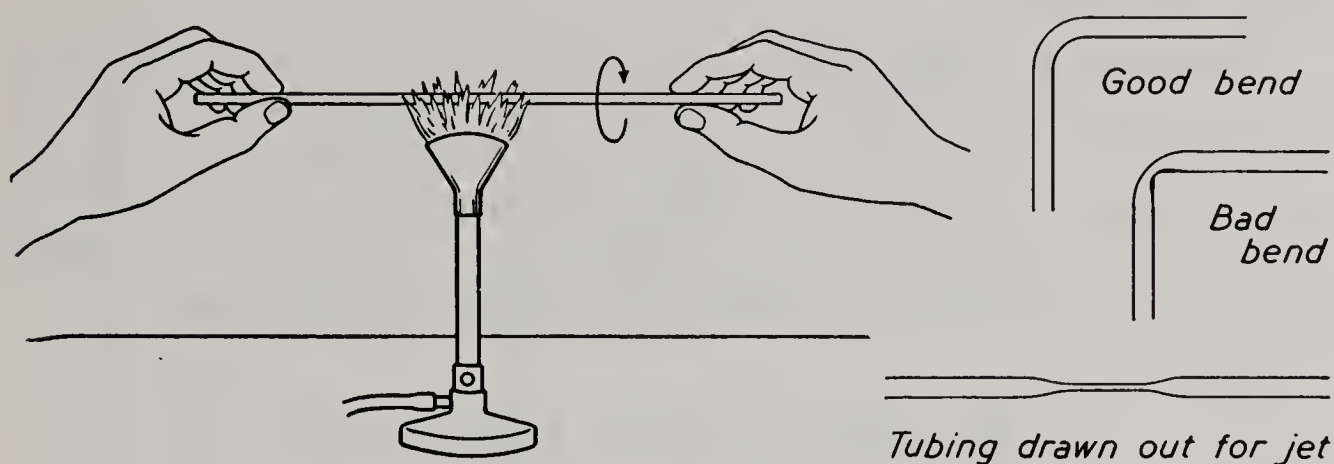


Fig. 12. Heating glass tubing.

Fig. 12. Continue the slow rotation until the tube is hot enough to sag of its own weight. Then *remove it from the flame* and immediately bend it to the desired angle. A block of wood or a similar right-angled object resting on an asbestos mat is of great value in achieving an exact right angle and in ensuring that all portions of the tubing will lie in the same plane. In cooling, the hot glass must not be allowed to come in contact with the surface of the desk, or any cold object.

If angles other than right-angles are desired, the bending is simplified if the desired angle is previously drawn in pencil on the asbestos mat. The heated tubing is then made to coincide with this angle.

Making a glass jet. Discard the "wing top" and heat the tubing in a non-luminous flame. Rotate the tube to insure uniform heating. When the heated region is soft and begins to sag, *remove from the flame* and slowly draw the ends apart in a straight line until the constricted portion is of the desired diameter. Allow the tubing to cool and then make a file scratch at the point where it is to be severed. Then fire-polish the cut end, being careful not to seal it.

Inserting glass tubing in stoppers. *Carelessness in this simple operation can easily result in a badly lacerated hand.* Be sure that the

ends of the tubing are properly fire-polished. Lubricate both the tubing and the opening in the stopper with water. Hold the stopper between the thumb and forefinger of one hand and grasp the *towel-wrapped* tube with the other hand about one inch from the end closest to the stopper. Slowly work the tubing into the opening by a gradual rotary motion.

Never grasp any bent tubing *at the bend* when inserting it into a stopper and never point tubing toward the palm of the hand which holds the stopper.

Constructing a wash bottle. A wash bottle is shown in Fig. 7. The nozzle should be drawn out so that only a moderately fine stream of water is forced out by gently blowing into the mouthpiece. All the ends of the glass tubing used must be fire-polished. Wash bottles are of value as containers for distilled water to be used in washing precipitates, making special solutions, rinsing glassware, etc.

Working with Balances

A. A Chemical Balance

The chemical balance is a delicate instrument that will weigh light objects to a high degree of accuracy and is of inestimable value to the chemist. Since the quantities of substances to be weighed usually are small, these balances have to be much more accurate than those in ordinary use in the commercial world. If one recalls that an ounce is equivalent to 28.35 grams and that in quantitative experiments one has to make accurate weighings in terms of tenths, hundredths, or even thousandths of a gram, it is apparent that this sensitive instrument must be used with a great deal of care, since it may be easily damaged by careless manipulation.

Fig. 13 represents an inexpensive simple model. A vertical pillar (**P**) is firmly attached to a wooden base. The metal beam (**B B₁**) acts as a lever and has a central, wedge-shaped knife-edge which points downward and rests upon a true surface or bearing-plane situated at the top of the pillar, when the balance is in the weighing position. This permits the beam to swing freely on the fulcrum with a minimum of friction. Some models have a V-shaped groove at the top of the pillar instead of the flattened bearing-plane. A knife-edge is also found at each end of the beam (**E₁** and **E₂**). These are turned upwards and support the stirrups (**S** and **S₁**) by fitting into V-shaped grooves with which the stirrups are provided. All the knife-edges and bearing-planes are constructed from polished agate

or hardened steel. Each stirrup has a notched hook which holds a pan support (**H** and **H₁**), which carry the removable scale-pans (**W** and **W₁**).

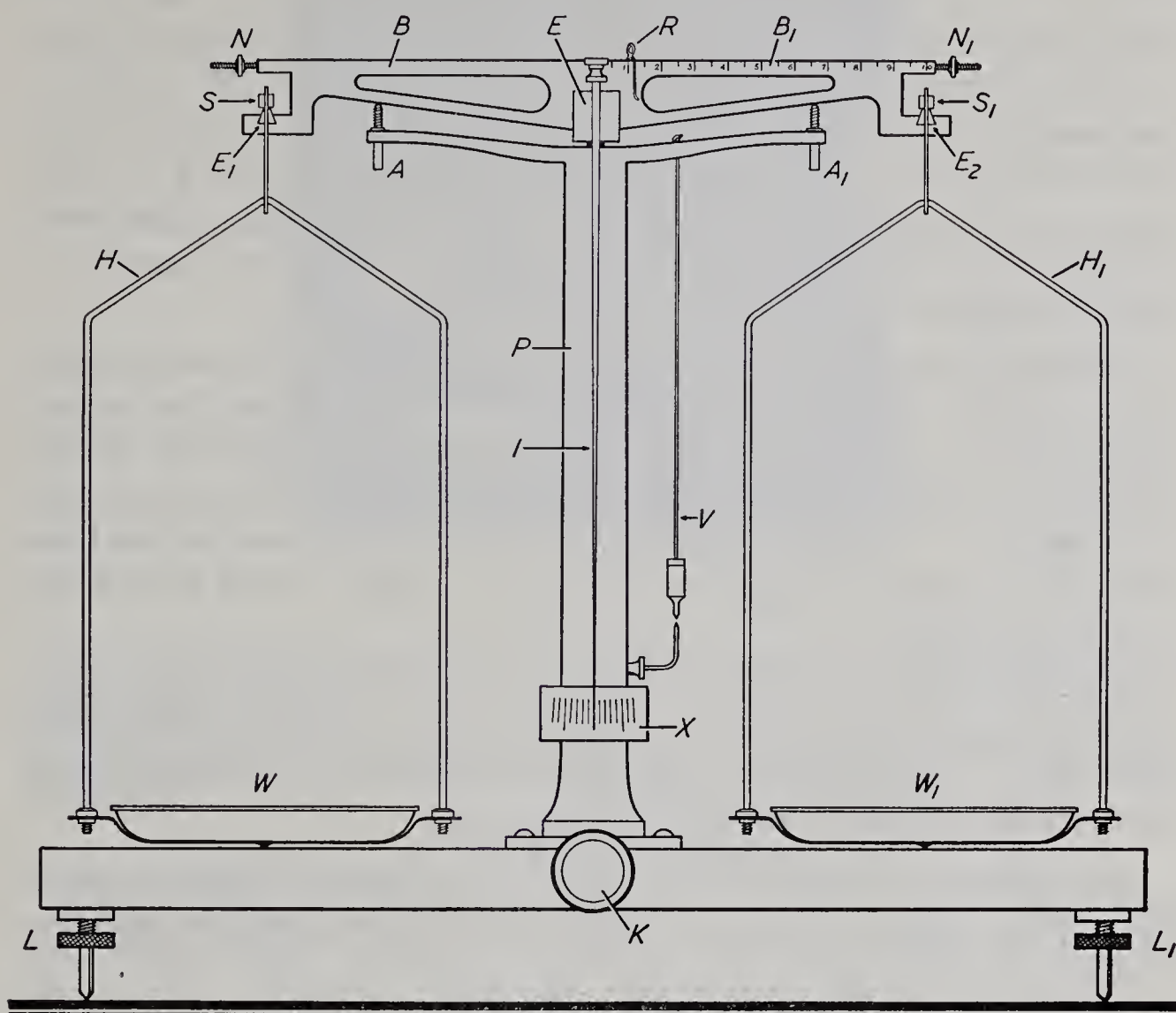


Fig. 13. A chemical balance.

When the balance is not in use, the beam is supported by two fixed arresting screws (**A** and **A₁**), thus relieving the central knife-edge of needless pressure and possible damage. With the beam supported by this arresting mechanism, the scale-pans rest on the wooden base and no stress is imposed on their particular knife-edges. The beam is raised and lowered by means of a knurled knob (**K**) located in the front of the balance. When turned to the right, the beam is lifted from the arresting screws and begins to swing, since it is now in the weighing position, resting on the central knife-edge. In this position, the stirrups are also suspended from their respective knife-edges. When the knob is turned in the opposite direction, the beam is lowered and comes to rest on the fixed screws which are intended to support it when not in use.

Attached to the beam is a pointer (**I**) which is free to move over a scale (**X**), thus indicating when a balance is attained between the object to be weighed and the weights used. If the two arms of the balance are in exact equilibrium, the pointer will swing equally on both sides of the middle of the scale. If the swing is incorrect it may be adjusted by turning one of the small balancing nuts (**N** and **N₁**) in the required direction. A faulty swing may also be due to the fact that the balance is not level. Two levelling-screws (**L** and **L₁**) on the base may be turned if a spirit level shows that the base is not level. A plumb-line (**V**) at the side of the pillar is sometimes used as a levelling device.

The sliding weight or rider (**R**) is a looped wire of known weight, bent so that it can be placed astride the beam. It may be moved along the beam when weights of less than .01 gm. are required. In the balance sketched, the rider may be moved along the right-hand side of the beam. In more expensive models, graduations occur on both sides of the centre, so that the weight of the rider can be subtracted as well as added.

A good balance should be housed in a cabinet with clear glass windows on the sides and top and fitted with a counterpoised sliding front door. This glass case is designed to exclude dust and laboratory fumes, and to avoid disturbing air currents.

The student is advised to sit in front of a chemical balance, read the description a second time, and identify the different parts by reference to Fig. 13.

Sets of metric weights. Examine the metric weights available in your laboratory and observe how they are marked and arranged. You will probably find:

(a) Brass weights (Grams): 100 gm., 50 gm., 20 gm., 20 gm., 10 gm., 5 gm., 2 gm., 2 gm., 1 gm.

(b) Aluminum or nickel fractional weights (Milligrams):

500 mg. (0.5 gm.)	10 mg. (0.01 gm.)
200 mg. (0.2 gm.)	5 mg. (0.005 gm.)
200 mg. (0.2 gm.)	2 mg. (0.002 gm.)
100 mg. (0.1 gm.)	2 mg. (0.002 gm.)
50 mg. (0.05 gm.)	1 mg. (0.001 gm.)
20 mg. (0.02 gm.)	
20 mg. (0.02 gm.)	

The box also contains a pair of forceps which must be used when weights are removed from, or replaced in, the box.



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Fig. 14. A box of metric weights.

You should check to ensure that all the weights are present and report any missing weights.

Make a list (or a diagram) showing the weights provided in your laboratory.

B. A Trial Weighing

Possibly your instructor will want you to do a practice weighing using some object such as a penny, a nickel, a dime, or a small beaker, etc. If such be the case, proceed as follows:

Check the swing. See that the stirrups are not displaced and that the scale-pans are clean and dry. Turn the operating knob, thus allowing the beam to swing, and observe the excursions made by the pointer. If the balance is in proper adjustment, the pointer will travel over an equal number of divisions on each side of the central graduation on the scale. If the pointer travels farther in one direction than in the other, return the beam to its position of rest by turning the operating knob, and adjust one of the small balancing nuts in the proper direction. Again observe the excursions of the pointer when the beam is raised. Repeat the adjustments in this manner until the swings to right and left are equal. It is not necessary to wait for the pointer to come to rest. A good balance will continue to swing for a long time.

Use the metric weights. *Never add or remove anything from the scale-pan while the beam is raised.*

Place the object on the *left*-hand pan. Select the largest weight in the box and *using forceps* place it on the *right*-hand pan. Now partially raise the beam by turning the operating knob. If the weight selected is too heavy, the scale-pan holding the object will rise and the pointer will swing to the left; a movement of the pointer to the right indicates that the weight is too light. Let us assume that the first weight selected is too heavy. Return the beam to its position of rest, remove the weight and return it to the box. Try the next largest weight, and continue in this manner from the largest to the smallest, without omitting any, until a weight is added which is lighter than the object. Leave this on the pan, add the next lighter weight, and so on through the brass weights. Then work through the fractional weights (milligrams) in the same systematic diminishing order of magnitude, until the pointer swings equally (or nearly so) in both directions.

If your balance is equipped with a rider, this is commonly used when weights of less than 0.01 gm. are required. This feature varies so widely in different models that the manipulation of the rider on your particular balance will be explained by the teacher.

The balance case should be closed when observing the movements of the pointer in the final stages of the weighing.

Record the weights. When you have completed the weighing, write down the separate weights carefully before they are removed from the scale-pan and total them at once. This record should be made in your notebook and never upon a loose sheet of paper, since the loss of the latter will necessitate a repetition of the weighing, and often the repetition of the entire experiment. Then, beginning with the largest weight, return them to the box and check the individual weights as you return them. In this way any error will be observed and corrected.

Special instructions. Most of the rules to be observed in the care and manipulation of the balance have been stressed at appropriate places in the earlier paragraphs. The following additional precautions must be observed:

1. Dust must be removed with a "camel's hair" brush.
2. The object to be weighed must be clean and dry.
3. Do not weigh an object when hot; the heat causes warm air currents which affect the weighing.
4. Chemicals are never placed on the unprotected scale-pan. Solids are placed upon a watch-glass or a piece of paper. Liquids are usually placed in a small beaker. The empty container is

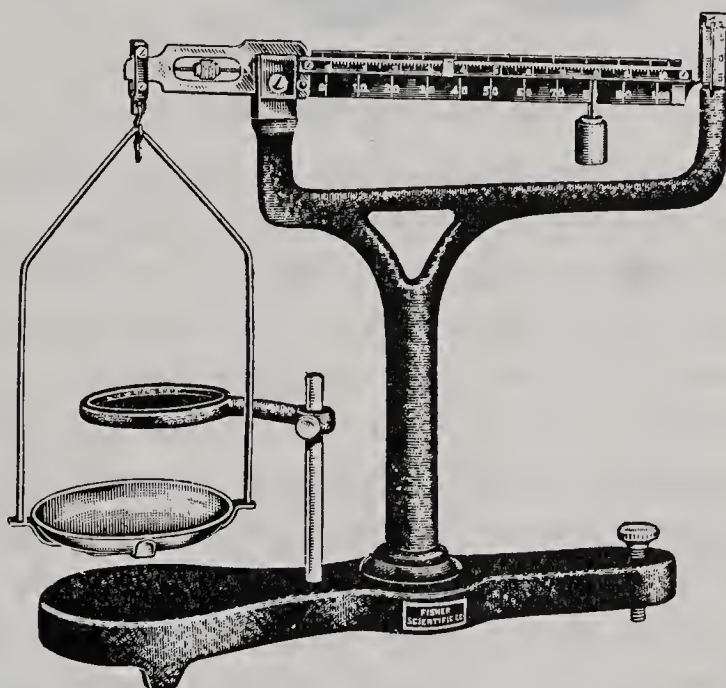
weighed and the weight of the substance is determined by subtraction.

5. Set the object to be weighed in the *centre* of the left-hand scale-pan. The larger weights should also be placed in a central position in their pan.

6. When a weighing is completed, be certain that the beam is left in the resting position and that the lid of the weight box is closed.

C. Triple Beam Balances

This type of balance has three separate graduated beams, mounted in parallel in the same horizontal plane and joined to an arm bearing the hardened steel knife-edge, stirrup, and adjusting thumb-screw. The front beam is graduated to 1 gram in tenths and hundredths of a gram divisions; the centre beam is graduated to 100 grams in 10 gram divisions; the rear beam is graduated to 10 grams in 1 gram divisions. The front beam has a sliding rider weight, while the other two are notched and are equipped with non-removable weights. A beam release mechanism controls the freedom of movement of the beam.



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Fig. 15. A triple beam balance.

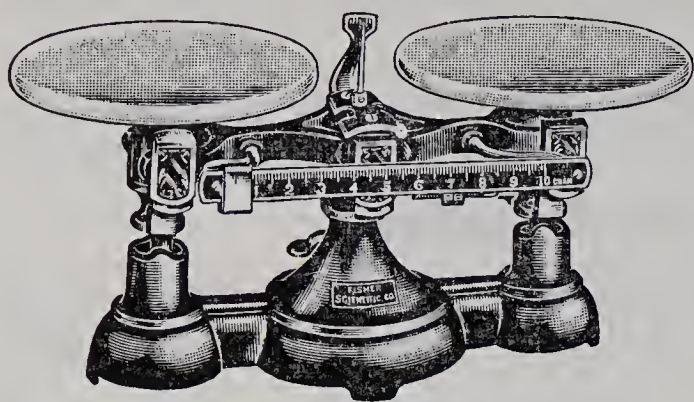
The base has an adjustable screw for levelling the balance and many models are equipped with a spirit level as well. A metal rod, attached to the base, acts as a holder for an adjustable beaker support for specific gravity determinations.

In some types, the free end of the beam is constructed to hold a weight hanger which is used when the load being weighed exceeds 111 grams. The weight of the hanger is 100 grams, and slotted auxiliary weights are provided to fit this hanger, thus increasing the weighing capacity.

D. Trip Balances

These are heavy-duty balances used when sensitivity is not of primary importance, as in dispensing chemicals, weighing large samples for making solutions, determining specific gravity, etc.

The two pans are non-corrosive vitrolite plates. The knife-edges are usually hardened steel with self-centering agate bearings. They



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Fig. 16. A trip balance.

usually have a single beam, graduated 0 to 10 grams in 0.1-gram divisions. A permanently attached rider eliminates the need for 10-gram and smaller weights. Some models are equipped with a second beam graduated 0 to 200 grams in 10-gram divisions, thereby extending the range over

which weighings may be made without accessory loose weights.

In using this type of balance, the student should be certain that the sliding rider is at the extreme left, or at zero. A light touch on one of the pans will set the balance swinging and the pointer should swing at least two or three divisions to either side of centre on the scale. If it swings farther to one side than to the other, a slight turning of the adjusting thumbscrew should be made until the pointer moves through as many divisions to the left as to the right. It is unnecessary to wait until the pointer comes to rest.

The object to be weighed must be placed on the *left* pan. Since jarring dulls the knife-edges and affects the sensitivity of the balance, this may be avoided by supporting one pan with the left hand when weights are being added or removed. The weight of the object is equal to the sum of all weights present on the right pan, plus the weight represented by any movement of the sliding rider from its zero position.

UNIT 1

A Study of Matter

EXPERIMENT 1. The measurement of volume.

1. Read the numbers on a graduated cylinder. Most graduates are marked in millilitres. A litre is defined as the volume occupied by one kilogram of air-free (boiled) distilled water at 4°C. This volume is 1000.027 cubic centimetres, so that 1 ml. is equal to 1.000027 cc. For practical purposes the millilitre and the cubic centimetre may be considered to be the same. Because most vessels used in the laboratory to measure the volumes of liquids or gases are graduated in millilitres, the millilitre will be used as the unit of volume throughout this manual.

What is the capacity of the graduated cylinder?

What does each division represent?

2. In measuring the volume of a liquid always keep the cylinder vertical, the eye level with the surface of the liquid, and read from the bottom of the meniscus (TECH. 1), except when measuring the volume of mercury in which case the top of the meniscus is the guiding line. Determine the capacities of several test-tubes of different sizes, a beaker, an evaporating dish, or any other containers in your locker, by filling them with water and measuring the volumes in a graduated cylinder. From the graduated cylinder pour 10 ml. of water into an empty test-tube. Stick a paper label on the tube, leaving a portion of the tube bare, and put a pencil mark opposite the water meniscus. You will frequently be asked to approximate certain volumes. For example, if asked to use approximately 5 ml. of a liquid it is not necessary to measure this volume precisely. You may add sufficient liquid to bring the level half way to the pencil mark. With practice you will learn to estimate different volumes fairly accurately.

Why is the eye always placed on a line horizontal to the surface of the liquid whose volume is being measured? Why is the graduated cylinder always placed on a table?

Why do you read from the bottom of the meniscus?

Record the volumes of the containers you have used.

3. Examine a burette. Support the burette on a retort stand by means of a burette clamp and partly fill it with water. Take a reading by observing the bottom of the meniscus. Record the reading.

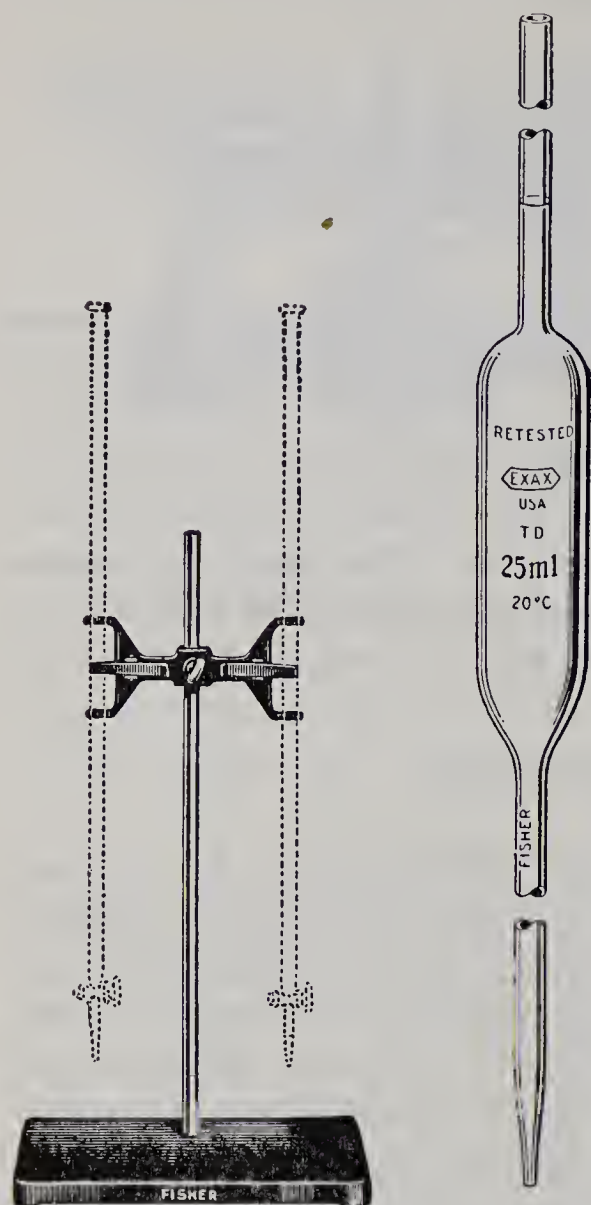
Does the reading you have taken represent the volume of water in the burette? Explain.

Do the graduations read from the bottom to the top or from the top to the bottom?

Place a beaker below the burette and allow some water to run into it. Read the burette and record the reading.

What volume of water is in the beaker?

4. Examine a pipette. Place the pointed end of the pipette in water and the opposite end in the mouth and by suction draw water into the pipette well above the mark on the stem. Then quickly place the first finger over the upper end of the pipette and at the same time withdraw the pipette from the mouth. If the water level drops below the mark on the pipette the operation must be repeated. Gently ease the finger and



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Fig. 17. Burette clamp supported on a stand and holding two burettes. On the right, a 25-ml. pipette.

allow the water to drip slowly until the bottom of the meniscus is even with the mark on the pipette. Now allow the water to flow into a beaker draining out the last drop by touching the tip of the pipette against the side of the beaker.

What is the volume of the pipette used? Find out the volumes of the pipettes that are available in your laboratory.

Why is there danger when a pipette is used to measure the volume of a poisonous liquid? How could you avoid this danger and still use a pipette?

QUESTIONS

1. Compare the uses of a graduated cylinder, a burette, and a pipette under the headings of (a) convenience, and (b) accuracy.
2. Make diagrams showing (a) a graduated cylinder containing 47.5 ml. of water, (b) a graduated cylinder containing 38 ml. of mercury, (c) a burette with the meniscus at the 60 ml. level, and (d) a 10 ml. pipette properly filled with water.

EXPERIMENT 2. The determination of density.

1. Weigh an empty beaker. Pour a carefully measured volume of water into the beaker, using either a graduated cylinder or a pipette, and weigh the beaker and the water.

What is the density of water in grams per cc.?

Describe how you measured the volume of the water.

2. Use a lump of roll sulphur about one half inch in diameter, but small enough to permit dropping it into a graduated cylinder. Weigh the sulphur and record its weight.

What is the weight, in grams, of the lump of sulphur?

3. Add water to a graduated cylinder until it is approximately half full of water and record the volume. Drop the piece of sulphur into the water and record the volume.

What is the difference in volume? What is the cause of this difference?

What is the volume of the sulphur in cubic centimetres?

Find the density of sulphur in grams per cc. What is the specific gravity of sulphur?

4. Repeat the above experiment using a piece of iron instead of sulphur.

What is the density of iron in grams per cc.? What is the specific gravity of iron?

QUESTIONS

1. If the iron used in the above experiment formed a perfect cylinder, describe how you could find the density of iron without using the method of displacement.
2. Describe how you could continue with Procedure 1 in the above experiment to get a more accurate value for the density of water.

✓ EXPERIMENT 3. **Changes in state.**

1. Heat a test-tube half full of water to the boiling point. Place a small piece of Wood's metal in the boiling water and observe its behaviour. Cool the water and note any changes in the metal.

Describe the changes that take place in the metal. Are these physical or chemical changes? What names are used to express these changes in state?

Suggest how you could find the melting point of Wood's metal.

State an industrial use of Wood's metal that is a direct application of a property learned from this experiment.

2. Place some crystals of naphthalene in a clean dry test-tube to a depth of one inch (TECH. 8). Half fill a beaker with water and set it on a wire gauze placed on a ring attached to a retort stand. Place the test-tube containing the naphthalene in the water and place a thermometer in the naphthalene so that the bulb is surrounded by the crystals. Heat the water until it becomes lukewarm and then lower the flame so that its temperature rises slowly. Observe the temperature at which the naphthalene melts. Continue heating until the naphthalene has completely liquefied and then remove it from the beaker of water. Note the temperature at which solidification begins.

At what temperature did the naphthalene melt?

What was the temperature at which solidification began? What is the melting point of naphthalene? Did the temperature change during solidification?

What two changes in state are illustrated in this experiment?

3. Recall experiments that you have performed in previous grades and answer the following questions.

Is it correct to say that the boiling point of water is 100°C.?

What condition must be recognized when reference is made to the boiling point of any liquid?

Define "boiling point".

What term is used to express the change in state from a gas to a liquid? What is the heat relationship in such a change?

✓ 4. Place a spoonful of naphthalene in an evaporating dish (Pyrex is preferable). Cover the dish with a filter paper and then invert another evaporating dish over the first one so that the filter paper separates the two dishes (Fig. 18). Heat the naphthalene gently

until it has disappeared from the lower dish. (If it should melt during the heating, reduce the temperature.) Examine the filter paper and the upper evaporating dish.

What happened to the naphthalene in the lower dish?

Describe what was observed on the filter paper and upper dish.

What must have been the state of the naphthalene to allow it to pass through the filter paper?

What change in state is illustrated in this experiment?

(TEXT p. 10)

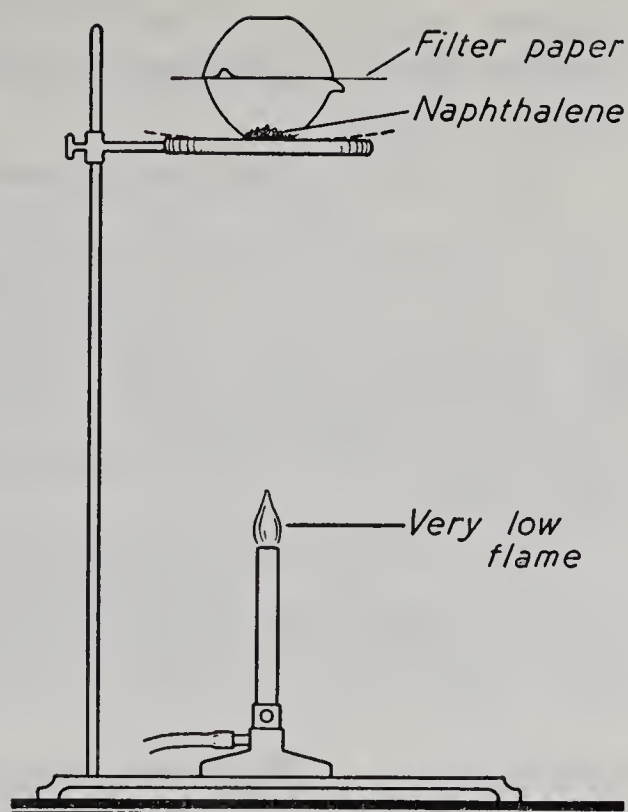


Fig. 18. The sublimation of naphthalene

QUESTIONS

1. Describe an experiment to show that if the pressure is reduced the boiling point of water is lowered.
2. Describe an experiment to show that if the pressure is increased the boiling point of water is increased.
3. Give an example from the above experiments for each of the changes in state shown in Fig. 2.1 in the text.
4. Give an example of a sublimation that occurs in nature and describe the conditions under which such a phenomenon occurs.

EXPERIMENT 4. Physical and chemical properties.

1. Examine a lump of roll sulphur and a piece of iron such as were used in Experiment 2. Wet a finger and touch it to the sulphur and then to the tongue. Repeat with the iron. Attempt to scratch each in turn with a knife. Attempt to break each substance into two pieces. Bring a magnet to touch each substance in turn. Shake a small piece of each substance separately in about 5 ml. of water in a test-tube, filter, collect a few drops of the filtrate on a watch-glass and evaporate to dryness (TECH. 5).

Tabulate your observations so as to make a comparison between

sulphur and iron in the following properties: colour, odour, taste, density (Expt. 2), hardness, brittleness, magnetic property, and solubility in water.

Why is each of the above properties called a physical property?

2. Put a small piece of sulphur about the size of a pea on a piece of mica placed across a ring supported on a retort stand, and direct the hot flame of a Bunsen burner on the sulphur. Observe any changes in colour, state and odour. Repeat with iron.

Did the sulphur undergo any permanent change? Give reasons for your answer.

What is the effect of heat on the iron?

What property of sulphur and what property of iron is illustrated by this experiment? Are these properties physical or chemical? Explain.

✓ **EXPERIMENT 5. The identification of a few common substances.**

1. Place samples of salt, sugar, starch, powdered zinc, and baking soda on separate pieces of paper and number the papers according to a prearranged plan. The substances selected are harmless and can be touched, tasted, and smelled without danger, but the identification of unknown substances by these methods is not recommended as a general practice.

Record the colour, odour, and taste of each substance examined.

2. Put separately about 5 gm. of each substance on a piece of mica placed across a ring supported on a retort stand, and turn the hot flame of a Bunsen burner directly on each substance in turn.

Does the substance burn? Is any permanent change observed?

3. Into separate test-tubes place a small quantity of each substance. Add a few drops of hydrochloric acid to each and observe.

Describe any evidence of a chemical reaction between the acid and the substance tested.

4. Into separate test-tubes place a small quantity of each substance. Add 5 ml. of distilled water to each and shake. Heat each to the boiling point and cool. Save each test-tube and its contents for Procedure 5.

Are the substances soluble in cold water?

What was the effect of heat in each case?

5. Divide the cold contents of each test-tube mentioned in Procedure 4 into two equal parts. To one portion add a few drops of

silver nitrate solution and to the other a few drops of iodine solution (tincture of iodine).

Record the effect of each reagent on each substance.

QUESTIONS

1. Collect the information learned about each substance and tabulate the properties of each under the headings: colour, odour, taste, solubility in water, inflammability, effect of hydrochloric acid, and the effect of adding (a) silver nitrate solution, and (b) iodine solution, to a solution of each substance.

2. From a comparison of the properties tabulated above, tell how you could distinguish each substance from the other four. Give as many differences as you can.

✓ EXPERIMENT 6. Some interesting chemical changes.

1. CAUTION: Do not inhale any gas produced.

Put a pellet or a few turnings of copper in a test-tube and add a few drops of concentrated nitric acid (TECH. 2).

List three observations.

What evidence is there that at least two substances other than the original copper and nitric acid are present?

2. Examine samples of lead nitrate and potassium iodide. To 5 ml. of a solution of lead nitrate in a test-tube add 5 ml. of a potassium iodide solution.

Describe the two substances used.

From the fact that you used solutions of these substances, what can be said about their solubilities?

What happened when solutions of the two substances were mixed? Give two observations.

Why would you conclude that at least one new substance has been formed?

3. On a piece of paper mix about 1 gm. of baking soda and 1 gm. of tartaric acid. Place the mixture in a test-tube and then add about 5 ml. of water.

What evidence is there of a chemical reaction?

4. (*Demonstration*) Mix one volume of powdered sugar with two volumes of potassium chlorate and place a half-teaspoonful of the mixture on a piece of mica supported by a ring attached to a stand.

Touch the mixture with a glass rod dipped in concentrated sulphuric acid.

What evidence is there of a chemical reaction?

✓ **EXPERIMENT 7. Heating metals in air.**

1. Clean a piece of copper wire with emery paper and carefully examine the cleaned surface. Using tongs hold the wire in the tip of a hot Bunsen flame and continue to heat it until no further change is noticeable. Allow the wire to cool and again carefully examine the surface. With a knife scrape off some of this surface coating and examine the flakes closely.

Describe the colour, lustre, and flexibility of the copper wire before heating.

Describe the colour, lustre, and flexibility of the flakes scraped from the surface of the wire. Can you distinguish two colours?

Explain the nature of the change that has taken place.

Clean the wire with emery paper once more, examine the cleaned surface, and repeat the heating.

Why is it necessary to clean the wire for the chemical change to continue?

2. CAUTION: To protect the desk surface, care should be taken to make sure that burning magnesium is never held above the unprotected surface.

Clean a piece of magnesium ribbon about three inches long with emery paper and carefully examine the cleaned surface. With tongs, hold the tip of the ribbon in a Bunsen flame. When it ignites hold the burning ribbon over a piece of clean glass and catch the product of combustion. Examine the new substance formed.

List four properties of magnesium.

Describe the action that takes place when magnesium is heated.

Describe the substance formed by the burning of magnesium and compare its properties with those of magnesium.

Has there been a physical or a chemical change? Give reasons for your answer.

3. Examine some clean steel wool. (Although not pure iron, steel wool closely resembles iron.) With tongs, hold the steel wool in a hot Bunsen flame for a few minutes. Remove it from the flame, allow it to cool and look for any new substance that has been formed.

Compare the properties of the new substance with those of iron.

What kind of change has taken place?

4. (*Demonstration*) Examine a clean platinum wire. Hold it in a hot Bunsen flame and observe any change in appearance while it is in the flame. Allow it to cool and again examine the platinum.

Was there any permanent change in the platinum?

QUESTIONS

1. Do all metals undergo a chemical change when they are heated in air?

2. Name two metals other than those used in this experiment that were heated in air by Lavoisier. (TEXT p. 16)

EXPERIMENT 8. Change in weight when a metal is heated in air.

1. Support a crucible with its cover on a clay triangle which in turn is supported by a ring attached to a stand (Fig. 19). Heat the crucible and cover gently at first and then, for at least one minute, with the hottest flame of your burner. Cool to room temperature and then weigh the crucible together with its cover.

2. Roll a piece of clean magnesium ribbon, about four inches in length, into a loose coil and place it in the crucible. Weigh the crucible, cover, and magnesium.

3. Return the crucible and contents to the clay triangle and place the inverted cover loosely over the crucible so that there is a narrow space through which air may enter. Heat, gently at first, and then more strongly, until all the magnesium has burned. It may be necessary (using tongs) to raise the cover from time to time to admit air to the crucible, but this must be done with care or some of the ash may escape. Cool and weigh the crucible, lid, and contents.

4. Record the weights as follows:

1. Weight of crucible and cover gm.
2. Weight of crucible, cover, and magnesium gm.
3. Weight of magnesium used (2—1) gm.
4. Weight of crucible, cover, and ash gm.
5. Weight of ash formed (4—1) gm.

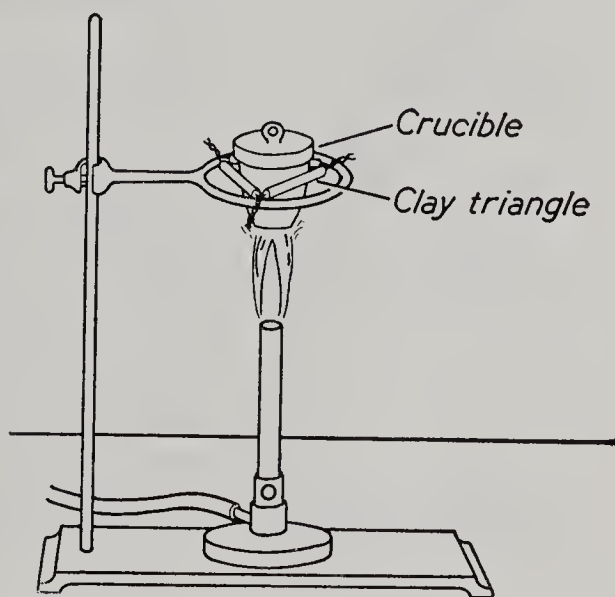


Fig. 19. Heating a covered crucible and its contents.

How much did the magnesium increase in weight when it was burned? What is the cause of this increase in weight? What percentage is this of the original weight?

5. Some of the class may substitute copper for magnesium in the above experiment. If copper is used, very fine wire or foil is suitable, and the crucible with the copper should be heated with a high flame for at least 15 minutes.

QUESTIONS

1. Account for the increase in weight experienced by some metals when they are heated in air.
2. Why was the crucible heated before it was weighed?

UNIT 2

Oxygen

EXPERIMENT 9. Heating mercuric oxide.

1. Place about as much mercuric oxide as would lie on a ten cent piece in a small Pyrex test-tube (TECH. 8). Hold the test-tube as in Fig. 20 and heat gently, keeping the end containing the solid in the flame. After heating for a very short time, remove the test-tube from the flame and allow it to cool.

List the changes in the appearance of the mercuric oxide when hot and when cold.

2. Continue to heat the mercuric oxide, increasing the temperature until it is being strongly heated. Lower a glowing splint of wood slowly into the test-tube until it almost touches the material in the bottom of the tube, and just as slowly withdraw it. For a convenient reference the accompanying diagram (Fig. 20) is numbered.

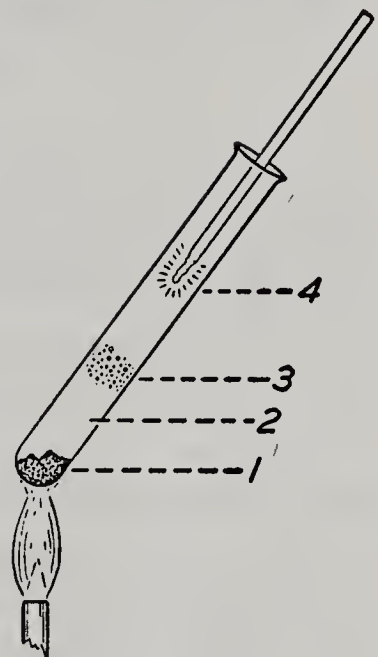


Fig. 20. Heating mercuric oxide.

Copy Fig. 20 in your notebook.

Describe and account for what happened to the glowing splint when it was placed in Position 4.

Describe the change in the condition of the splint as it was slowly lowered to Position 2.

Give two reasons for concluding that the mercuric oxide was undergoing a chemical change.

Describe the change in the splint when it was again raised to Position 4.

Continue heating the test-tube until the mercuric oxide in the bottom of the tube has disappeared. Remove the tube from the flame and allow it to cool. Examine the sides of the tube.

Describe the colour and lustre of the material found at Position 3. What is its physical state?

Name the substance present at Position 3.

Describe and identify the substance seen just below Position 3.

Reconsider the observations made during the experiment when the glowing splint was raised and lowered in the test-tube while the mercuric oxide was being strongly heated.

Account for the change in the condition of the splint as it was moved from 4 to 2 and also for the change when it was moved from 2 to 4.

What is the factor that caused the change in the rate of burning of the splint?

Name the two substances that are present in the gaseous state at 2.

Which of these condenses at 3? Why?

Why does the other substance not condense?

QUESTIONS

1. Write a word equation to represent the chemical reaction that takes place when mercuric oxide is heated.

2. Who was the first chemist to prepare oxygen from mercuric oxide? What was the name by which this chemist knew mercuric oxide? What was the name he used for oxygen? (TEXT p. 15)

3. When mercuric oxide is heated, **decomposition** takes place. When mercuric oxide is formed the action is called **combination**. Compare the temperatures at 1 and 3 (Fig. 20) and draw a conclusion as to which is the lower temperature, the temperature of decomposition or the temperature of combination.

4. Would mercuric oxide be a suitable source of oxygen in quantity in the laboratory? Give reasons for your answer.

EXPERIMENT 10. The effect of heat on some substances that contain oxygen.

Put about one-half inch of each of the following substances (each substance listed contains oxygen) into separate Pyrex test-tubes: potassium permanganate, sand, calcium oxide, lead dioxide, potassium nitrate, sodium nitrate, copper oxide, and potassium chlorate. Heat each test-tube, at first gently and then more intensely. Test each with a glowing splint.

Do all substances that contain oxygen liberate oxygen when they are heated?

From which of the above substances is oxygen liberated by heating?

List these substances in the order of the ease with which they liberate oxygen.

EXPERIMENT 11. **Oxygen from potassium chlorate and the effect of manganese dioxide as a catalyst.**

✓ 1. Add crystals of potassium chlorate to a dry Pyrex test-tube until there is a depth of about one-half inch and heat gently. Gradually increase the temperature until the potassium chlorate melts and then bring a glowing splint to the mouth of the tube. Continue heating until bubbles begin to come off. Again bring the glowing splint to the mouth of the test-tube.

What is the cause of the crackling sound heard when crystals of potassium chlorate are first heated? (TEXT p. 22 and p. 76)

Is there any evidence that oxygen is being given off from the potassium chlorate just at the time it is melting?

What gas is coming from the molten bubbling potassium chlorate?

Continue the heating and test from time to time with a glowing splint until the splint will not rekindle. The substance left in the test-tube is called potassium chloride.

What change in state is noted in the substance in the test-tube?

Why did the bubbling cease?

Which has the lower melting point, potassium chlorate or potassium chloride? Give a reason for your answer.

✓ 2. Heat the same quantity of potassium chlorate as used in Procedure 1 above, until it just melts. Drop in a pinch of manganese dioxide and test with a glowing splint.

Note any change in the liquid and the effect on a glowing splint held at the mouth of the tube.

Compare the temperature at which oxygen was liberated in Procedure 1 with that in Procedure 2. What are the effects of adding the manganese dioxide?

✓ 3. Mix together on a piece of paper approximately the same quantity of potassium chlorate as used in Procedure 1, and approximately one-third this quantity of manganese dioxide. Place this mixture in a Pyrex test-tube. In another Pyrex test-tube place an equal quantity of

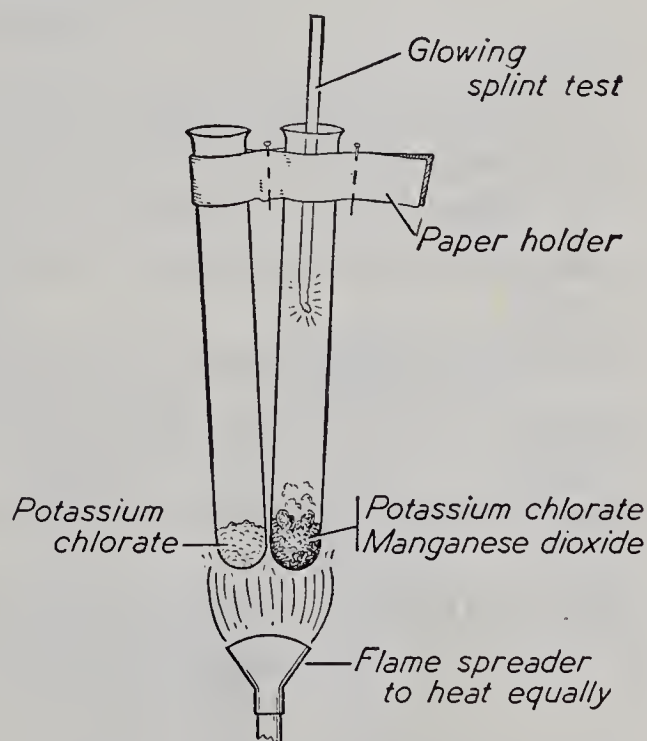


Fig. 21. The effect of heating manganese dioxide with potassium chlorate.

potassium chlorate only. Hold both test-tubes in a common holder and, using a flame spreader, heat the two uniformly under the same conditions (Fig. 21). Test each tube at frequent intervals with a glowing splint.

Set aside the test-tube containing the manganese dioxide for use in Procedure 4.

From which tube is the oxygen liberated first?

From which is it liberated more rapidly?

What is the effect of having manganese dioxide present?

4. Add water to half fill the test-tube saved from Procedure 3 and heat until near boiling. Filter (TECH. 4).

Describe the residue on the filter paper.

What substances are present in the filtrate?

5. Dry the filter paper containing the residue from Procedure 4 and scrape a small amount of the dried residue from the paper. Put some potassium chlorate crystals to a depth of about one-half inch in a Pyrex test-tube. Heat the tube until the potassium chlorate just melts and then add a pinch of the dried residue. Test with a glowing splint.

What was observed when the residue was dropped into the molten potassium chlorate? What gas was liberated?

What property has this residue in common with manganese dioxide?

In what other respects does the residue resemble manganese dioxide?

QUESTIONS

1. Write a word equation expressing the reaction that takes place when potassium chlorate is strongly heated.

2. What name is given to express the type of reaction resulting from the use of manganese dioxide in the above experiment? (TEXT p. 23)

3. What conclusions can be made regarding the action of manganese dioxide on potassium chlorate in the production of oxygen? Suggest an experiment designed to determine if any manganese dioxide has been used up in this reaction.

EXPERIMENT 12. (*Demonstration*) Weight relations when manganese dioxide is used as a catalyst with potassium chlorate in the preparation of oxygen.

1. Put about 1 gram of manganese dioxide on a filter paper and accurately weigh the two together. Save the filter paper for future

use. Mix the manganese dioxide with approximately three times its volume of potassium chlorate and put the mixture in a Pyrex test-tube. Heat the mixture until all bubbling ceases and then continue to heat for about five minutes.

2. When the test-tube is cool, half fill it with distilled water and bring the liquid nearly to the boiling point. Filter the liquid using the filter paper on which the manganese dioxide was originally weighed. To remove all of the material from the test-tube add distilled water from time to time, each time pouring the liquid into the filter paper. Continue to wash the test-tube in this manner until all of the solid is removed. Thoroughly wash the residue on the filter paper with distilled water (TECH. 9). Allow the residue and filter paper to dry completely.

3. Weigh the filter paper with the attached residue.

What is the residue?

Is there any change in the weight of the manganese dioxide?

What additional information does this experiment give about the action of the manganese dioxide?

Why was the residue so thoroughly washed?



EXPERIMENT 13. To prepare, collect, and study the properties of oxygen.

PREPARATION AND COLLECTION

1. Thoroughly mix on a piece of paper a teaspoonful of potassium chlorate with approximately a third of a teaspoonful of manganese dioxide and fill a Pyrex test-tube one quarter of its depth with the

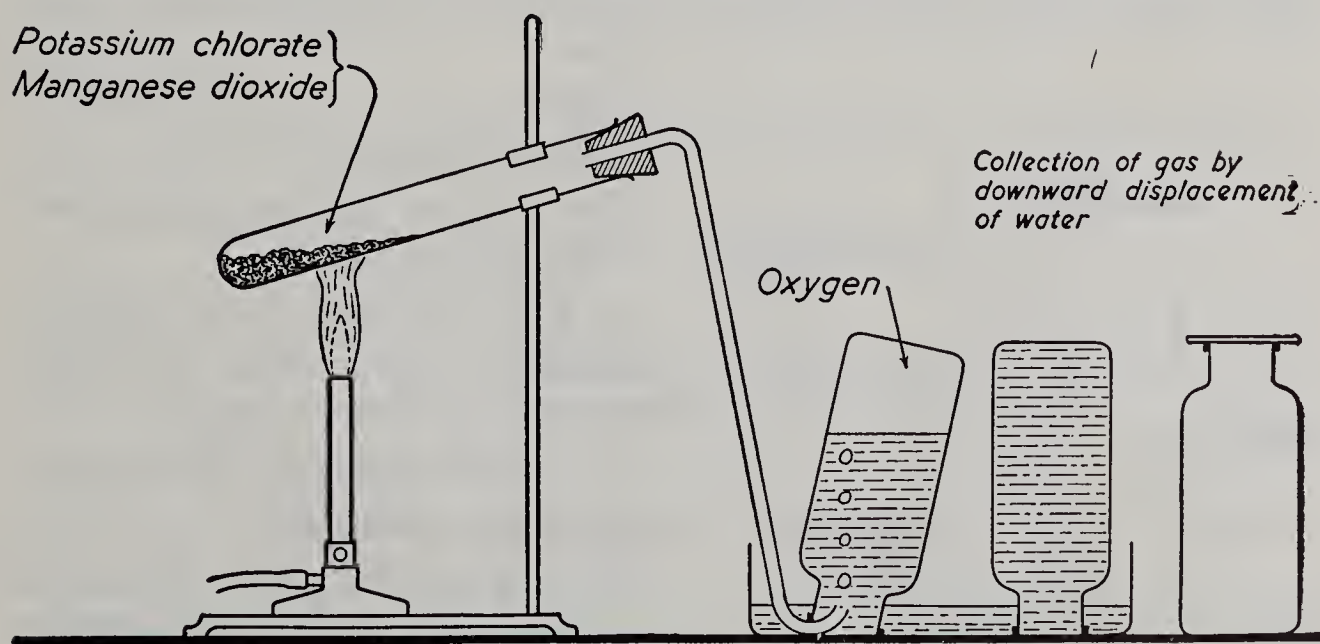


Fig. 22. The laboratory preparation and collection of oxygen.

mixture. Distribute the mixture along the sides of the lower half of the tube.

2. Set up the apparatus as shown in Fig. 22. Before the gas bottles are placed in the pneumatic trough, fill each with water until the rounded upper surface of the water extends above the mouth. Slide a glass plate over the mouth of the bottle and hold it in this position until the bottle is inverted in the water in the pneumatic trough, and then remove the glass plate (TECH. 12).

3. Begin heating the mixture nearest the mouth of the tube and gradually work towards the base. Heat gently so as to obtain a slow but steady stream of bubbles in the pneumatic trough. Reject the first bottle of gas collected. *Why?*

4. When each bottle is full, cover the mouth with a glass plate and lift it out of the water, placing it mouth upward on the desk. When the last bottle has been collected lift the delivery tube out of the water in the trough before removing the burner. *Why?*

(Note: More than one laboratory period will be required to complete the work that follows. It is suggested that some of the following be demonstration experiments.)

PROPERTIES OF OXYGEN

1. Examine a bottle of oxygen.

Note the colour and odour of the gas. What evidence have you that oxygen is not very soluble in water?

2. Insert a glowing splint into a bottle of oxygen. Extinguish the flame and repeat the experiment several times, using the same bottle of gas.

Compare the rate of reaction in oxygen with the rate in air.

Does oxygen support combustion?

Does oxygen burn?

3. Pour approximately 10 ml. of limewater into a bottle of oxygen and shake for half a minute.

Is there any change in the appearance of the limewater?

Heat a piece of wood charcoal in a deflagrating spoon (or hold with tongs) until it glows and then lower

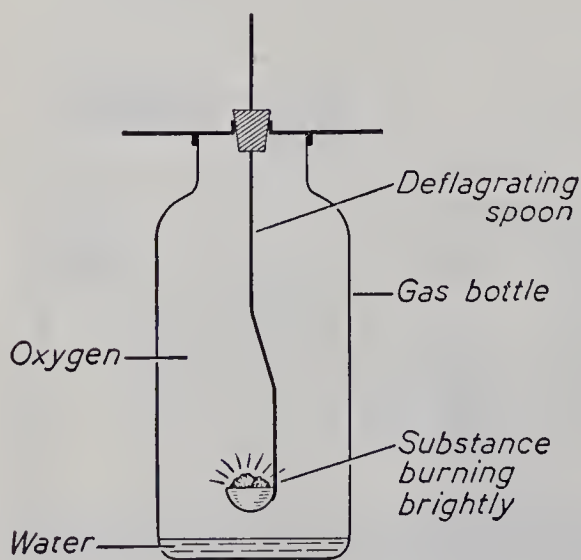


Fig. 23. Burning substances in oxygen.

it quickly into the bottle of oxygen containing the limewater. Allow the charcoal to remain there until no further action takes place and then remove it.

Compare the burning of charcoal in air and in oxygen.

What happens to the limewater?

What gas was formed in the bottle? (TEXT p. 27)

Write a word equation for the burning of charcoal.

4. Lower a piece of glowing charcoal into another bottle of oxygen and hold it there until no further action takes place. Then remove it. Pour about 10 ml. of water into the bottle and shake vigorously for half a minute. Pour half of this liquid into a beaker.

Place a small piece of red litmus paper and a small piece of blue litmus paper in the liquid remaining in the bottle and leave for several minutes. If you are in doubt as to whether or not there has been any change, put small pieces of the two litmus papers into distilled water, which does not affect litmus, and compare with the colours of the litmus in the liquid being tested.

Describe any change in the colour of either of the litmus papers.

Add a few drops of neutral (green) bromthymol blue solution to the liquid in the beaker.

Note any change in colour.

Which is the better indicator in this experiment, litmus or bromthymol blue? Why?

What property of the solution is shown by the use of these indicators? (TEXT p. 28)

5. Line a deflagrating spoon with asbestos paper and place in it a piece of roll sulphur about the size of a bean. Ignite the sulphur in a Bunsen flame.

Note the colour, size, and brightness of the flame and any change in state or colour of the sulphur.

Lower the burning sulphur into a bottle of oxygen.

Note any changes in the flame.

After the flame is extinguished, remove the spoon from the bottle.

Is there any unburned sulphur remaining in the spoon?

Why did the flame go out?

Smell the gas remaining in the bottle, using your hand to waft a little of the gas towards your nose.

Describe the odour.

What is the name of this gas? (TEXT p. 28)

Write a word equation for the burning of sulphur in oxygen.

Add about 10 ml. of water to the bottle, shake for half a minute, and test the liquid with litmus papers (as in Procedure 4).

What do the results indicate about the nature of this liquid?

Name this solution. (TEXT p. 29)

Write the word equation that represents the formation of this new substance.

6. Wind a piece of magnesium ribbon, about three inches long, into a spiral by twisting it around a match, or a piece of glass tubing. Remove the spiralled ribbon and hold it in a pair of tongs. Ignite the tip of the spiral in a Bunsen flame.

What is observed just before the magnesium ignites?

Lower the burning magnesium into a bottle of oxygen, being careful not to let it contact the sides of the bottle.

Compare the burning of magnesium in air and in oxygen.

Describe the nature of the oxide found in the bottle and name it.

Add about 10 ml. of water to the bottle, shake and test the liquid with red and blue litmus papers.

Does any of the material dissolve?

What do your observations indicate about the nature of this liquid?

Name the substance present in the solution. (TEXT p. 29)

Write a word equation representing the action that took place when the product of combustion was shaken with water.

- ✓ 7. (*Demonstration*) Place a pinch of red phosphorus in an asbestos-lined deflagrating spoon. Ignite the phosphorus and quickly introduce the spoon into a bottle of oxygen.

Compare the burning of phosphorus in air and in oxygen.

Why is the burning of phosphorus more rapid in oxygen than in air?

Describe the product of the combustion of phosphorus in oxygen.

Phosphorus forms two oxides (TEXT p. 141). When there is a high concentration of oxygen, phosphoric oxide is formed. As the oxygen concentration diminishes, phosphorous oxide is produced. In this experiment, the product of combustion is probably a mixture of these two oxides. *Why?*

Add about 10 ml. of water to the bottle, shake, and test with litmus.

Classify the liquid as acidic or basic.

Write a word equation for the burning of phosphorus in oxygen.

Write a word equation to show the reaction when the product of com-

bustion was shaken with water. (You may regard it as chiefly a solution of phosphoric acid.)

(Note: If yellow phosphorus is used, the result is more spectacular, but great care is necessary. This form of phosphorus is kept under water and must be cut under water. It must be handled with forceps; never with the hands. Any unburned phosphorus remaining on the spoon should be destroyed by complete burning. Even a trace of yellow phosphorus exposed to the air could be the source of a major fire.)

✓ 8. (*Demonstration*) Place a small piece of sodium in an asbestos-lined deflagrating spoon and heat it until it ignites.

Describe any changes in the sodium before it ignites.

Does it ignite easily?

Compare the ignition temperatures of sodium and phosphorus.

What is the colour of the flame?

Lower the burning sodium into a bottle of oxygen.

Describe the burning in oxygen and the product of combustion.

Could the product contain a mixture of oxides? (TEXT p. 241)

Shake the product of combustion with water and test with litmus.

Classify the oxide(s) of sodium.

Write word equations to represent these reactions.

✓ 9. (*Demonstration*) Add sufficient water to a bottle containing oxygen to occupy a depth of one inch. Using tongs, heat a piece of fine iron wire (or steel wool) until it glows brightly. Quickly lower it into the bottle of oxygen.

Record your observations.

Why was water added to the bottle before the introduction of the iron?

Compare the colour and brittleness of the product of combustion (magnetic iron oxide) with that of the original iron.

Shake some of the oxide with water and test with litmus.

Does magnetic iron oxide appear to be soluble in water?

Classify this oxide. (TEXT p. 28)

✓ 10. To a bottle filled with oxygen add 5 ml. of a solution of pyrogalllic acid and 5 ml. of a solution of potassium hydroxide. Place the hand firmly over the mouth of the bottle and shake.

What is the effect of oxygen on a mixture of these two solutions?

What effect was noticed on the hand? What does this indicate?

The above reaction is used as a conclusive test for oxygen. (TEXT p. 28)

QUESTIONS

1. Make a neatly labelled diagram of the apparatus used to prepare and collect oxygen.
2. Name three methods used in the collection of a gas. (TEXT p. 24) Which method was selected for the collection of oxygen? Why?
3. Summarize the results of Experiment 13 by completing the following table in your notebook:

ELEMENT BURNED	METAL OR NON-METAL	NAME OF OXIDE	DESCRIPTION OF OXIDE	EFFECT OF AQUEOUS SOLUTION ON LITMUS	CLASS OF OXIDE
Carbon					
Sulphur					
Phosphorus					
Magnesium					
Sodium					
Iron					

4. Name, define, and give an example of each of the three classes of oxides.
5. What type of element gives rise to (a) an acidic oxide, (b) a basic oxide? By what other terms are oxides known? (TEXT p. 28)
6. What is produced when an acidic oxide reacts with water? Give an example.
7. What is produced when a basic oxide reacts with water? Give an example.

✓ EXPERIMENT 14. The effect of adding water to sodium peroxide.

1. Clamp a test-tube to a stand and add sodium peroxide to the tube to a depth of one-half inch. Add a few drops of water and feel the bottom of the tube. Test at the mouth of the tube with a glowing splint.

Is the reaction between sodium peroxide and water endothermic or exothermic in nature?

What gas is produced by this reaction?

2. Add enough water to half fill the test-tube and when all action ceases test the solution with neutral litmus paper.

What is a second product of the reaction between sodium peroxide and water?

Write a word equation for the reaction between sodium peroxide and water.

Would sodium peroxide be a suitable source of oxygen in quantity?

✓ EXPERIMENT 15. (*Demonstration*) Ignition or kindling temperatures.

1. CAUTION: Do not touch the phosphorus with the hands (TECH. 10).

Using filter paper, dry a piece of yellow phosphorus about the size of a grain of wheat and place it on a piece of mica or asbestos paper. Put a piece of sulphur of approximately the same size near it. Heat a glass rod, not warmer than can be held with the hand, and touch it to the sulphur and then to the phosphorus.

Describe what happens.

Which has the lower ignition temperature, sulphur or phosphorus? Why is the precaution given about handling yellow phosphorus?

2. Examine a "strike anywhere" match. The tip is phosphorus sesquisulphide, a substance with a very low ignition temperature.

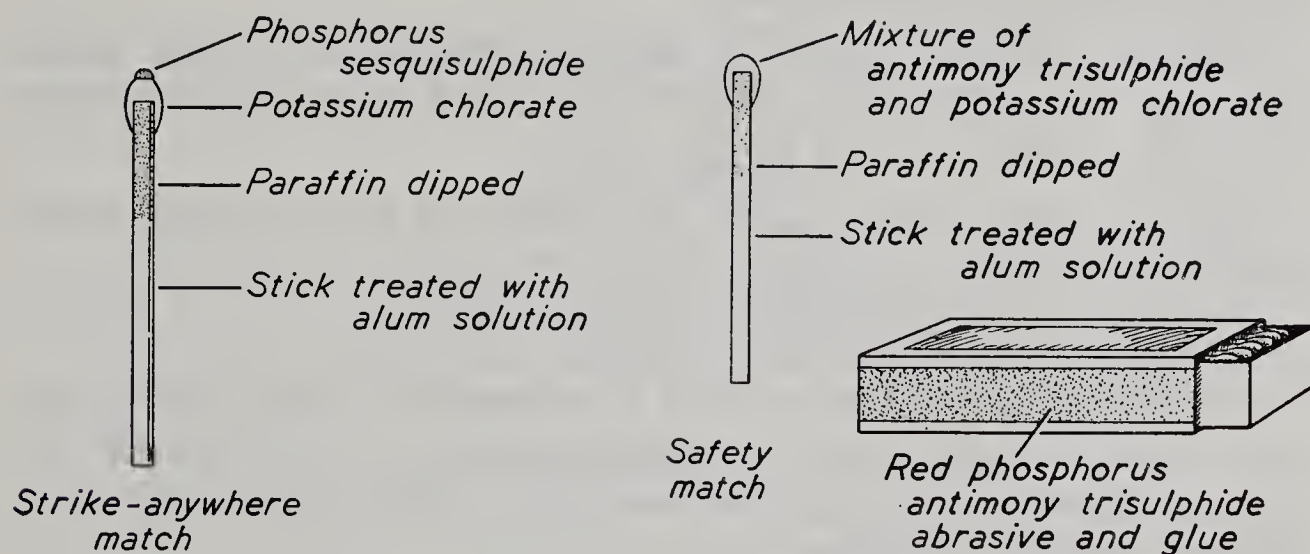


Fig. 24. "Strike anywhere" and "safety" matches.

The coloured material surrounding the tip is chiefly¹ potassium chlorate. Rub the match on a piece of fine sandpaper.

Make a diagram of the match and label all the parts.

Describe and explain what happens when the match is struck on the sandpaper. What is the purpose of (a) the phosphorus sesquisulphide, and (b) the potassium chlorate?

Note any change in appearance in the wood near the flame. Explain.

What two effects has an increased temperature on paraffin?

The wood of a match has been dipped in a water solution of alum or sodium sulphate or some other suitable chemical, and then dried. The wood is so treated to prevent smoldering. Light a strike-anywhere match and allow it to burn about half way down. Blow it out.

Why does blowing extinguish the flame?

Is there an after-glow?

Ignite a wooden splint. Allow it to burn for a short time and then blow it out.

Is there an after-glow? Explain.

3. The tip of a "safety" match usually consists of a mixture of antimony trisulphide, potassium chlorate, and glue. The material on the side of the match container (packet or box) is a mixture of red phosphorus, antimony trisulphide and powdered glass, held together with glue. Rub a safety match on a piece of fine sandpaper and then rub it on the material provided for striking.

Make a diagram of a safety match and label all the parts.

What happens when the match is rubbed (a) on sandpaper, and (b) on the special material provided?

Friction changes a trace of the red phosphorus to the yellow allotropic form (TEXT p. 172), which has a much lower ignition temperature than the red form.

Explain why a safety match will ignite on the relatively smooth surface provided for striking.

Was there any after-glow in the safety match?

4. (*Demonstration*) Put one-half a teaspoonful of corn starch on a piece of mica placed across a ring supported on a retort stand, and turn the Bunsen flame on it for some time. When the starch starts to burn, remove the flame.

Does the starch continue to burn?

Has starch a low or high ignition temperature?

Put another half teaspoonful of starch in a paper cylinder made by rolling a piece of foolscap to about one-half inch in diameter. Hold one end of the paper tube about four inches from a Bunsen flame and blow the starch from the tube into the flame.

Describe and explain what happens.

In comparing the ignition temperatures of two substances, why is it necessary to have conditions alike in all respects?

5. (*Demonstration*) Punch a hole in an eight-pound honey pail about one inch from the bottom. Bend a piece of metal tubing about five inches long to form a right angle, and solder a small metal funnel on one end. Push the metal tubing through the hole in the pail and solder it in position as shown in Fig. 25. Place a candle so that the

wick is about an inch above the top of the funnel and attach a piece of rubber tubing about two feet long to the metal tubing.

Put a half-teaspoonful of lyco-
podium powder in the funnel. Light the candle and, leaving the cover off the pail, give a short quick puff into the open end of the rubber tubing and observe.

Repeat the experiment but this time, after lighting the candle, push the cover on the pail and immediately give a quick puff into the open end of the tube and observe.

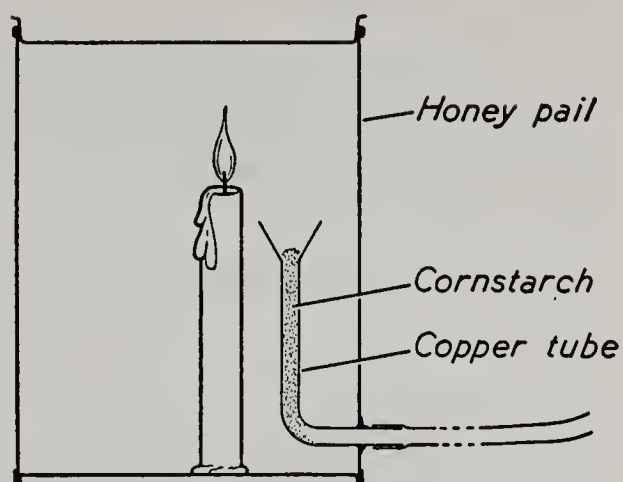


Fig. 25. A dust explosion.

The experiment may be repeated using corn starch or finely sifted flour.

Explain why a flash fire resulted when the pail was not covered and an explosion occurred when the cover was on the pail.

QUESTIONS

1. Arrange the substances phosphorus sesquisulphide, wood, and paraffin in the order of their ignition temperatures.

2. Why does phosphorus sesquisulphide ignite when rubbed over a rough surface?

3. Why is yellow phosphorus not used in matches?

4. Which has the higher ignition temperature: phosphorus sesquisulphide or antimony trisulphide? Give a reason for your answer.

5. Why does a large wooden stick not take fire as quickly as a small one?

6. Explain why dust explosions sometimes occur in grain elevators.

EXPERIMENT 16. Spontaneous combustion.

1. (*Demonstration*) Dissolve a piece of yellow phosphorus about the size of a small pea in approximately 5 ml. of carbon disulphide. Pour the solution on a piece of filter paper placed in the sink. Remove the filter paper from the sink, place it on a wire gauze and let it stand for a few minutes.

Describe and account for the phenomenon observed.

2. (*Demonstration*) Place a ball of absorbent cotton about two inches in diameter on a wire gauze supported on a ring stand. Flatten

the top of the cotton and spread a teaspoonful of sodium peroxide in a layer over it. Add a few drops of water to the centre of the sodium peroxide.

Describe and account for the phenomenon observed.

Why should sodium peroxide never be left unguarded in the laboratory or in an uncovered bottle?

UNIT 3

Air

EXPERIMENT 17. (*Demonstration*) The preparation of impure nitrogen from air.

1. Hollow out a slight depression in the centre of a flat cork. Coat this surface with a mixture of plaster of Paris and water and allow the mixture to harden. Float the cork on the surface of water in a pneumatic trough or the sink. The water should be about one inch in depth.

2. **CAUTION:** Do not touch the phosphorus with the hands (TECH. 10).

Place a small piece of yellow phosphorus about the size of a grain of wheat in the depression in the cork. Warm a piece of iron wire in the Bunsen flame and touch the phosphorus with the hot end of the wire. Quickly cover the cork and burning phosphorus with a wide-mouth bottle in such a way as to seal the mouth of the bottle beneath the water. Keep the bottle in place until all action ceases.

Describe what is observed.

Account for all the changes that take place inside the bottle.

Why does the phosphorus stop burning before it is all consumed?

Name the gas left in the bottle and account for its presence there.

3. Place a glass plate over the mouth of the bottle. Remove the bottle and place it in an upright position on the table. Remove the glass plate and quickly insert a blazing splint. Place another blazing splint in a bottle containing air.

Compare the behaviour of the burning splint in nitrogen and in air.

QUESTIONS

1. List four impurities in the nitrogen secured by this method. (TEXT p. 34 and p. 38) What is the approximate percentage of the impurities?

2. Suggest how the above experiment could be modified so that the proportion by volume of oxygen and nitrogen in air could be determined.

✓ **EXPERIMENT 18. To find the proportions by volume of oxygen and nitrogen in air (using steel wool).**

1. Degrease some steel wool by dipping it several times in gasoline. After the gasoline has evaporated moisten the steel with water and push it into a large-mouth bottle (Fig. 26). Invert the bottle in a dish of water and allow it to stand for several days. Record observations from time to time.

2. After the water stops rising in the bottle, adjust the level of the water in the dish until it coincides with the level of the water in the bottle. When this position is attained, place a cover glass over the mouth of the bottle and remove the bottle and its water contents from the dish.

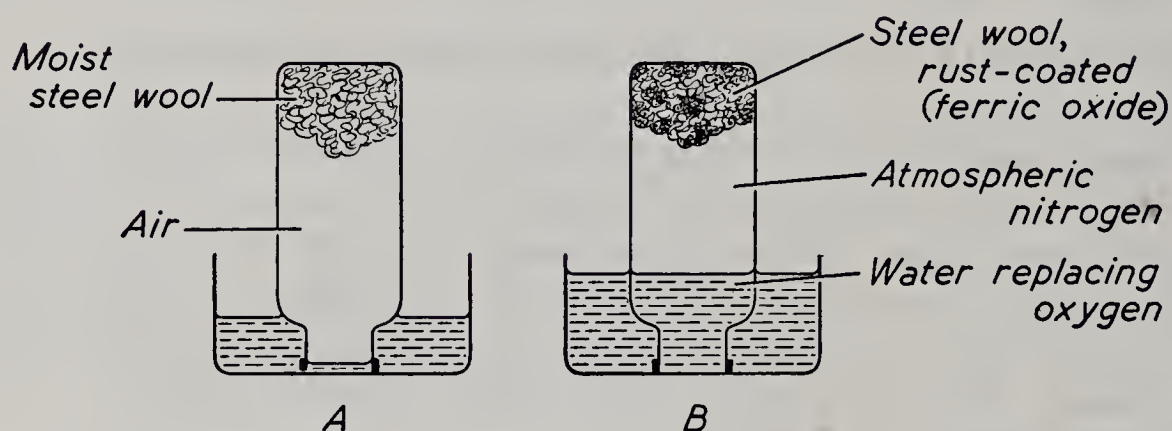


Fig. 26. The rusting of iron in air.

3. Use a graduated cylinder to measure the volume of the water in the bottle.

Record the volume of water that came into the bottle. What gas and what gas volume does this represent?

Fill the bottle with water and measure the volume of water it contains.

Record the volume of water needed to fill the bottle.

What gas and what gas volume does this represent?

Describe any change observed in the steel wool. What is the significance of the fact that all the iron has not rusted?

Calculate the percentage by volume of oxygen and of nitrogen in air.

QUESTIONS

1. Why was the water level adjustment made in the above experiment before removing the bottle from the dish?

2. Point out all the possible sources of error in the above experiment.

3. Compare your calculation of the percentage of oxygen in the air with the percentage shown in the TEXT, p. 37. What is your percentage of error? (See page 3.)

EXPERIMENT 19. (*Demonstration*) To find the proportions by volume of oxygen and nitrogen in air (using pyrogallic acid and potassium hydroxide).

1. Pour approximately 60 ml. of a concentrated pyrogallic acid solution into a 200 ml. gas-measuring tube (Fig. 27). Slant the tube and slide a stick of potassium hydroxide about an inch long (or a dozen small pellets) into it, quickly insert a stopper, and holding the stopper in place, invert the tube.

Record the volume of air enclosed in the tube.

2. Shake the mixture back and forth in the tube for about ten minutes.

What changes take place in the mixture in the tube?

3. Invert the tube in a deep vessel of water and, with the stoppered end well below the surface of the water, remove the stopper.

Describe and account for what happens.

What gas remains in the tube?

4. Cool to room temperature and adjust the level of the liquid inside the tube to coincide with the level of the water in the vessel.

Record the volume of gas remaining in the tube. How could this gas be identified?

Why was the adjustment in levels made before recording this volume?

Calculate the percentage by volume of oxygen and nitrogen in air.

What is your percentage of error?

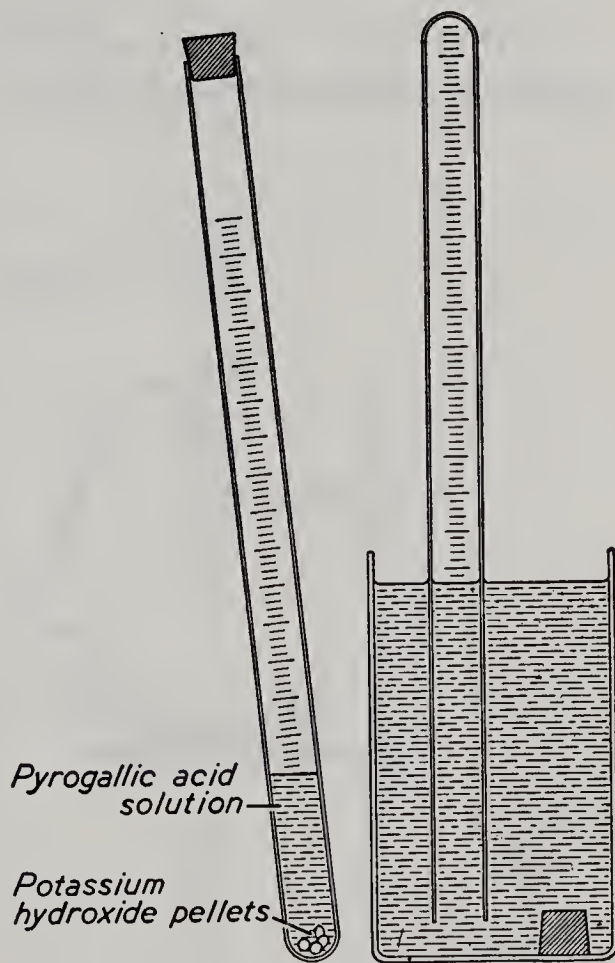


Fig. 27. Determining the composition of air by volume.

EXPERIMENT 20. (*Demonstration*) To prepare, collect and study the properties of pure nitrogen.

1. Prepare a solution of sodium nitrite by stirring 20 gm. of the salt in 100 ml. of water. Place 30 ml. of a saturated solution of ammonium chloride in a flask fitted with a dropping funnel and a delivery tube (Fig. 28). Place the flask in a water bath and pour some of the sodium nitrite solution into the dropping funnel. Have a beaker of cold water handy.

CAUTION: A violent explosion may occur in the generator if the reaction becomes too vigorous.

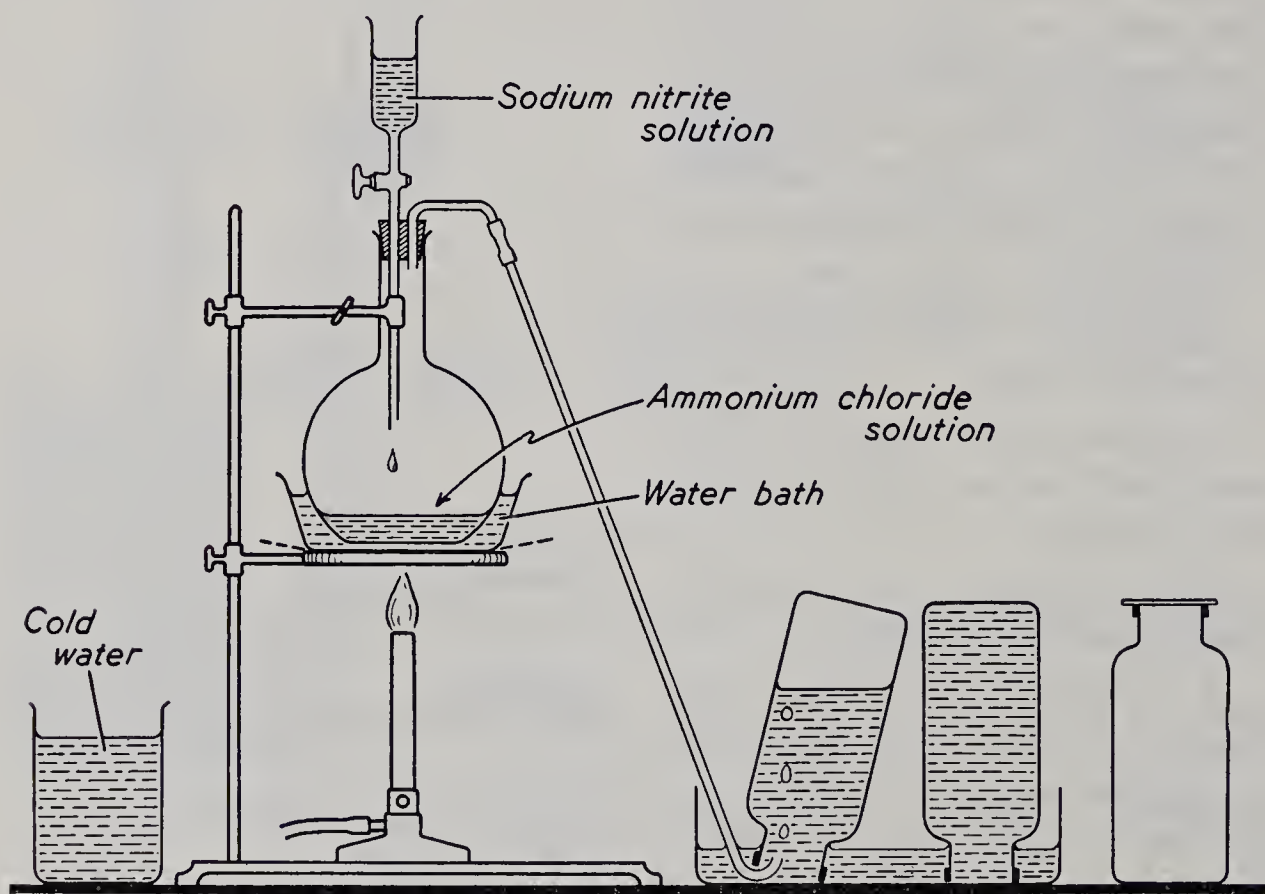


Fig. 28. The laboratory preparation and collection of pure nitrogen.

2. Heat the water in the water bath until warm and begin adding the solution of sodium nitrite from the dropping funnel, drop by drop. Should the contents of the flask froth considerably, remove the Bunsen flame and add cold water to the water bath so that the temperature is quickly reduced.

Collect three bottles of gas by the downward displacement of water. Discard the first bottle collected. *Why?*

3. Place the remaining two bottles in an upright position on the table using a glass plate to retain the nitrogen. Insert a burning splint into one of the bottles of nitrogen, and burning sulphur in a deflagrating spoon into the other.

List four physical properties of nitrogen.

What two chemical properties of nitrogen are indicated by using the burning splint and burning sulphur?

QUESTIONS

1. Write the word equation for the preparation of nitrogen.
(TEXT p. 34)
2. Why is it not advisable to prepare pure nitrogen by heating ammonium nitrite?

UNIT 4

Hydrogen

✓ EXPERIMENT 21. The action of sodium and potassium on water.

1. Remove a piece of sodium from the bottle in which it has been stored and place it on a piece of glass.

What is the liquid in which sodium is stored?

Why is sodium kept in this liquid?

Describe the appearance of the piece of sodium.

2. With a sharp knife cut a slice of sodium approximately one-eighth of an inch thick and then cut this slice into small cubes.

What is the appearance of the freshly cut surface of sodium?

How would you describe the hardness of the metal?

Describe the changes that take place in the appearance of the surface of the sodium following exposure to air. Account for these changes.

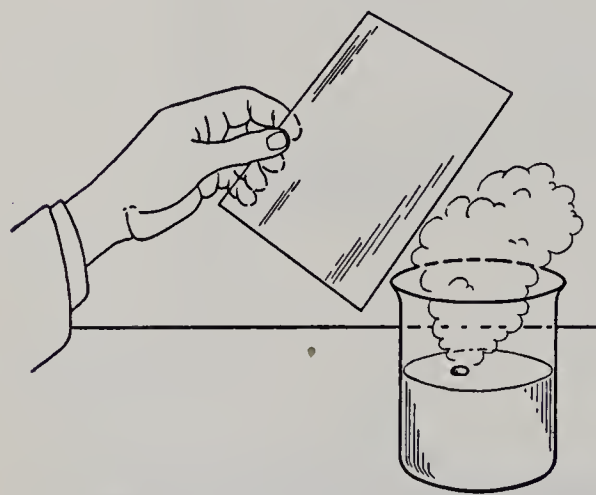


Fig. 29. Using a glass plate in watching the action of sodium or potassium on water.

3. Pick up a piece of sodium with a pair of forceps and place it on a piece of filter paper. Using another piece of filter paper gently press the sodium to remove any adhering liquid. Again pick up the sodium with the forceps and drop it into a beaker half-filled with water. Immediately cover the beaker with a glass plate to protect the eyes should the sodium explode. (For a demonstration experiment a large glass pneumatic trough is recommended.)

What change takes place in the shape of the sodium when placed on water?

What other changes take place in the appearance of the sodium?

What is the position of the sodium relative to the surface of the water?

Describe the motion of the sodium and account for the sound produced.

Describe the substance formed where the sodium is in contact with air.

Describe the substance formed where the sodium is in contact with the water.

What evidence is there that the action of sodium on water is exothermic?

4. Loosely wrap a piece of dry sodium in copper foil or lead foil and pierce the foil with the point of a knife blade several times. Drop the weighted sodium into a beaker of water and observe.

Describe the action that takes place.

Why is it necessary to wrap the sodium in the foil?

5. Repeat Procedure 4 and by placing inverted test-tubes full of water over the weighted sodium, collect (a) a test-tube full of the gas, and (b) a test-tube one-quarter full of the gas. Remove each test-tube from the beaker, and allow the water in the partly filled tube to be replaced by air. Keep the tubes in an inverted position and bring a lighted splint to the mouth of each test-tube in turn.

What is the colour of the gas in the test-tube?

Describe the action that takes place when the lighted splint is brought (a) to the test-tube full of the gas, and (b) the test-tube where the gas is mixed with air.

(The gas produced by the action of sodium on water is hydrogen and the burning of the gas as shown in 5(a) and 5(b) constitutes a characteristic test for hydrogen gas.)

6. Add a piece of neutral (purple) litmus paper to the solution remaining in the beaker.

Describe the colour change in the litmus.

What substance must be present in solution?

Account for the production of this substance.

7. Evaporate to dryness in a test-tube about 1 ml. of the solution remaining in the beaker.

What is the appearance of the residue in the test-tube?

What is the name of the residue? (TEXT p. 44)

8. (*Demonstration*) In an evaporating dish containing about 5 ml. of water place a small piece of filter paper folded so as to form a small pocket. Drop a piece of sodium into the water contained in this restricted area.

CAUTION: The sodium may explode.

Describe the action of the sodium.

What happens to the water?

Account for (a) the flame, (b) the flame colour.

9. (*Demonstration*) Using potassium instead of sodium, repeat Procedures 1 to 7, but omit Procedure 8. Record your observations.

CAUTION: When using potassium great care must be taken not to have the face too near the globule of potassium. When the flame seems to disappear, the molten globule explodes and showers the immediate area with molten metal and strong alkali.

QUESTIONS

1. In what respects are sodium and potassium similar?
2. In what respects do sodium and potassium differ?
3. Which element is more active? Give reasons for your answer.
4. What is the source of the hydrogen obtained in the above experiments?
5. Write word equations to represent the action of (a) sodium on water, (b) potassium on water.
6. List the properties of hydrogen discovered as a result of these experiments.

✓ EXPERIMENT 22. The action of calcium on water.

1. Half fill a 250 ml. beaker with water and invert four test-tubes filled with water into the beaker. Examine a small piece of calcium. Moisten a finger and place it on the calcium. Drop the calcium into the water in the beaker. Place the test-tubes in succession over the reacting calcium and collect (a) two test-tubes full of gas, (b) one test-tube half-full, and (c) one test-tube one-third full. Lift each test-tube from the beaker and stand it mouth downward on the top of the table. Save the contents of the beaker for Procedure 3.

Describe the appearance of the calcium.

Compare calcium with sodium (Expt. 21) as to (a) hardness, (b) lustre, (c) density, and (d) activity with water.

Describe the appearance of the contents of the beaker.

Is there any evidence of an exothermic reaction?

2. Leaving one test-tube full of the gas on the table, lift the others, one at a time, and keeping the mouth down, bring a lighted splint to the open end of each tube. Hold the remaining test-tube full of gas with the mouth up for one minute and then bring a lighted splint to the mouth.

What characteristic property enables you to identify the gas in each case?

What is the reason for keeping the test-tube's mouth down?

What happens to the gas in the test-tube held mouth up? Explain.

3. Add a piece of neutral litmus paper to the contents of the beaker (Procedure 1) and then filter some of the mixture (TECH. 4). Collect approximately one-half of a test-tube full of the filtrate in one test-tube, and one-quarter of a test-tube full in another test-tube. Place one end of a clean glass tube under the liquid in the test-tube half filled with the filtrate and blow your breath through the glass tube for several minutes. With a low flame, heat the tube one-quarter full of the filtrate until no liquid remains in the tube (TECH. 7).

Describe and explain the colour change of the neutral litmus.

What substance must be in solution to cause this colour change in litmus?

Account for the milky appearance of the contents of the beaker.

Describe the changes that take place when the breath is blown through the filtrate.

What is the common name for this solution?

Describe the appearance of the solid dissolved in the filtrate.

What is the chemical name of this solid?

4. Heat about 50 ml. of water in a beaker until the water boils. Remove the flame and when the water stops boiling drop a small piece of calcium into the hot water.

How does the rate of this reaction compare with the rate of the reaction of calcium on cold water?

QUESTIONS

1. In the above experiments, what evidence indicates that the hydrogen comes from the water?

2. From your observations, arrange the elements hydrogen, sodium, potassium, and calcium in the order of their relative activities. Do your experiments confirm the order as given in the TEXT p. 45?

3. Write a word equation to express the action of calcium on water.

✓ EXPERIMENT 23. The action of magnesium on steam.

1. Half fill two beakers with water and heat one to the boiling point. Into each beaker place a piece of clean magnesium ribbon and a piece of neutral litmus paper.

2. Repeat Procedure 1 using magnesium powder.

Describe the action of the magnesium ribbon on (a) cold water, (b) hot water.

Describe the action of the magnesium powder on (a) cold water, (b) hot water.

Do any of these experiments warrant collecting and testing the gas produced?

What conclusion can be drawn from the litmus test?

What effect was produced on the rate of reaction by (a) powdering the magnesium, (b) increasing the temperature?

3. (*Demonstration*) Half fill a 500 ml. flask with tap water and boil the water long enough to expel the dissolved air from the water. Wind a piece of magnesium ribbon approximately six inches long into a tight spiral small enough to be lowered easily into the mouth of the flask. Attach the spiral to the bowl of a deflagrating spoon or pick it up with suitable forceps or tongs and, after igniting the lower end of the spiral, carefully lower the burning magnesium into the steam-filled flask. Closely observe the reaction that takes place.

What do you observe about the magnesium the instant before it starts to burn in air?

Does the magnesium continue to burn when surrounded by steam?

Is there any evidence of a gas burning just above the mouth of the flask? If so describe the flame.

4. After the magnesium ceases burning, keep a small amount of the ash from falling into the flask and compare its appearance with the ash formed when magnesium burns in air. Add a piece of neutral litmus paper to the mixture of water and ash in the flask. Filter a small amount of the mixture in the flask and collect some of the filtrate in a test-tube. Evaporate the filtrate to dryness.

How does the product formed by the burning of magnesium in steam compare with the product formed by the burning of magnesium in air?

What does the litmus test indicate concerning the solution?

What is the chemical name of the ash produced in this experiment?

Describe the residue remaining in the test-tube. What is its chemical name?

QUESTIONS

1. What evidence is there to prove that hydrogen is produced when magnesium acts on steam?

2. What evidence is there to prove that magnesium oxide is produced when magnesium acts on steam?

3. Name two factors that affect the rate of a chemical reaction.

4. Compare the reaction rate of magnesium with (a) water at 100°C ., (b) steam at 100°C . What is the reason for the difference?

5. As a result of this experiment place magnesium in its proper position in the activity series.

6. Suggest a method of collecting the hydrogen produced in an experiment using magnesium and steam.

7. Write word equations to represent (a) the action of magnesium on steam, (b) the action of magnesium oxide on water.

✓ EXPERIMENT 24. (*Demonstration*) **The preparation of hydrogen by the action of iron on steam.**

1. Arrange the apparatus as shown in Fig. 30 and heat the finely divided iron as intensely as possible while the steam is passing through the combustion tube. Collect a gas bottle full of gas and test with a blazing splint.

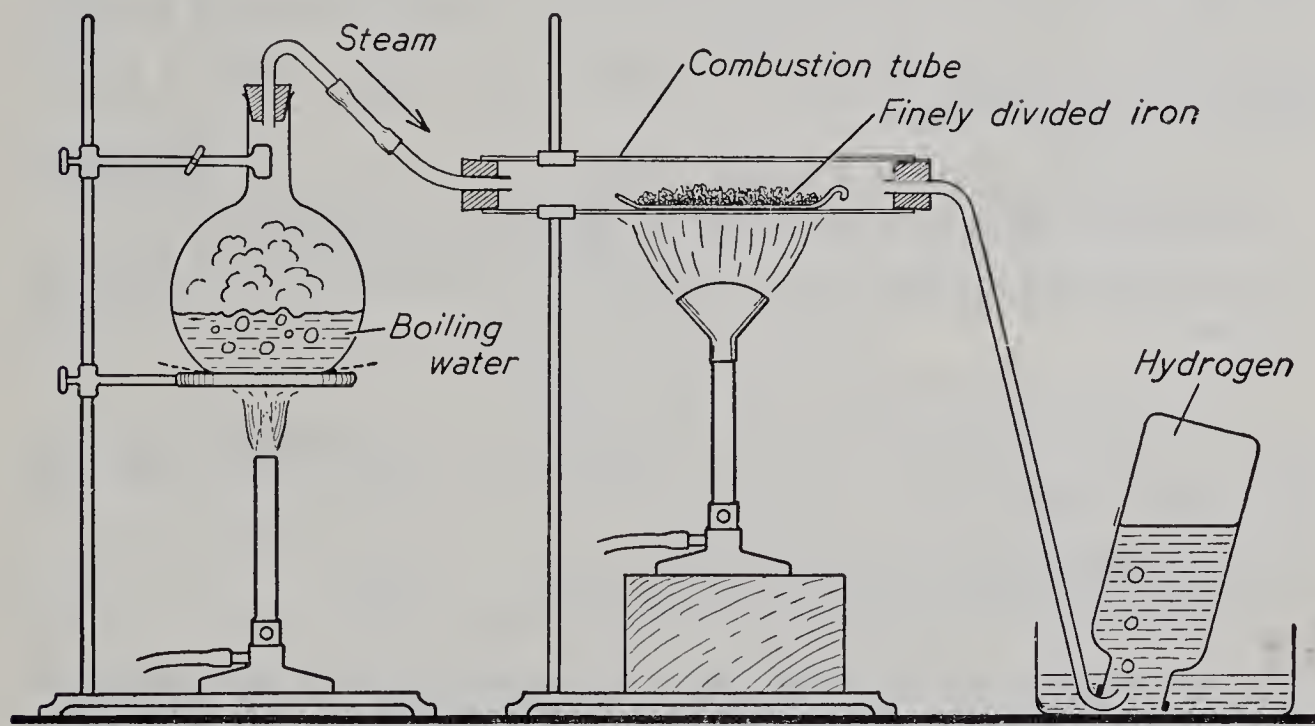


Fig. 30. Liberating hydrogen by the action of heated iron on steam.

- a) *Is there any evident chemical reaction within the combustion tube?*
- b) *Is the gas produced hydrogen? Explain.*
- c) *Is there any change in the appearance of the iron?*

2. To compare the action of iron on steam with that of magnesium on steam, replace the porcelain boat containing iron with one containing small pieces of magnesium ribbon, and collect and test the gas produced.

How does the activity of iron on steam compare with the activity of magnesium on steam?

Using evidence gained from this experiment, add iron to the activity series, placing it in its proper position.

QUESTIONS

1. Write the word equation to represent the action of iron on steam.

2. Why is this method of producing hydrogen used to a greater extent in industry than it is in the laboratory?

✓ EXPERIMENT 25. The action on dilute acids of metals above hydrogen in the activity series.

1. Into five separate test-tubes place (a) a small piece of calcium, (b) some magnesium ribbon, (c) a small piece of mossy zinc, (d) an iron tack, and (e) a small amount of iron filings. Add a small amount (4 ml.) of dilute sulphuric acid to each test-tube in turn. Note the reaction and test the escaping gas with a lighted splint. Record all observations.

Is a gas given off in each case? Where a gas is given off, identify it. Is there any difference in the rate at which the gas is given off?

To what factors would you attribute the difference in the rates of reaction?

What must be the source of the gas? Why?

2. Repeat Procedure 1 using (a) dilute hydrochloric acid, (b) dilute nitric acid.

Is hydrogen given off in each case?

Explain the action with nitric acid? (TEXT p. 47)

3. Into separate test-tubes place (a) a small piece of pure zinc, (b) a small piece of impure (mossy) zinc. To each test-tube add a few millilitres of dilute sulphuric acid.

Describe the action that takes place in each test-tube.

To what could the difference be attributed?

4. To the test-tube containing the pure zinc (PROCEDURE 3) add a small crystal of blue vitriol (cupric sulphate hydrate).

What effect has the blue vitriol crystal on this reaction?

Name and account for the small brown deposit remaining in the tube after the reaction.

Can blue vitriol be considered a catalyst in this reaction? Explain your answer.

5. To a small piece of pure zinc in a test-tube containing a small crystal of blue vitriol, add one ml. of concentrated sulphuric acid. Observe closely.

CAUTION: When water and concentrated sulphuric acid are mixed, so much heat is liberated that if the volume of water is small relative to the volume of concentrated acid, the heat converts the water into steam causing an explosion. The only safe way to dilute the acid is to add slowly the concentrated acid to the water. However, in this particular experiment, the following procedure may be followed if sufficient caution is exercised.

With a pipette or an eye dropper, allow *one drop* of water to fall into the tube and after all action ceases add one drop more. Repeat the addition of water, one drop at a time, and after each addition closely observe the reaction in the test-tube.

Is any hydrogen produced with pure zinc and concentrated sulphuric acid?

What happens when the first drop of water is added to the concentrated acid?

What is the effect of the gradual addition of water on the rate at which hydrogen is produced? Explain.

Explain why it is extremely dangerous to add more than one drop of water at a time to concentrated sulphuric acid.

6. Permit the action in Procedure 5 to continue for a short time, then filter the mixture and collect a small quantity of the filtrate (a) in an evaporating dish, (b) in a test-tube. Heat the evaporating dish strongly at first, but when the volume of liquid is considerably reduced, evaporate to dryness with a much smaller flame (TECH. 5).

CAUTION: As the solution becomes more concentrated small pockets of steam cause the liquid to spatter.

Put the test-tube containing the filtrate in your locker for a few days.

Why is it important to control the heating in this evaporation?

What is the colour of the residue? What is the chemical name for this substance? Account for its formation.

Describe the appearance of the contents of the test-tube after standing for a few days. Explain.

QUESTIONS

1. From information gained from these experiments suggest the most easily controlled method of obtaining large volumes of hydrogen from dilute acids.
2. Write word equations representing the action of calcium, magnesium, zinc, and iron on the dilute acids used.
3. Do the results of Experiment 25 confirm the position in the activity series of the metals used?
4. Suggest a reason for not using sodium and potassium with dilute acids in the production of hydrogen.

✓ EXPERIMENT 26. To prepare, collect, and study the properties of hydrogen.

PREPARATION AND COLLECTION

CAUTION: Do not strike a match or create a flame of any kind until so instructed.

1. Set up the apparatus as shown in Fig. 31 and place a handful of mossy zinc in the flask. Add dilute sulphuric acid through the thistle-tube until the zinc is well covered, making sure that the lower end of the thistle-tube is covered by the acid.

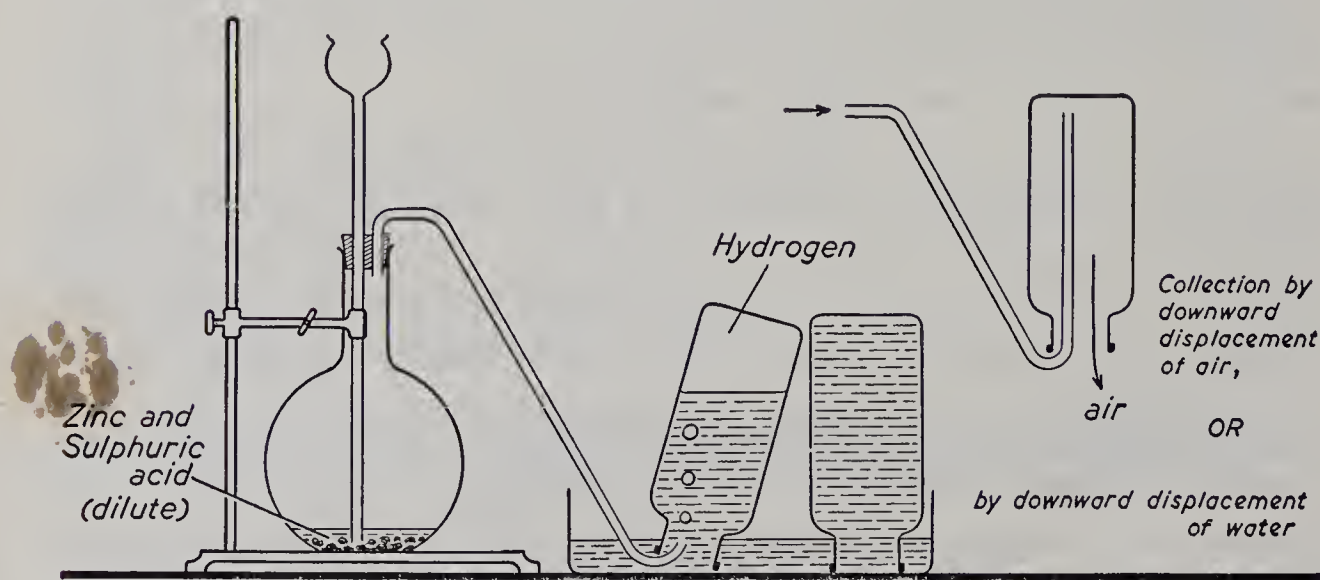


Fig. 31. The laboratory preparation and collection of hydrogen.

2. Let the reaction proceed for a short time and then collect, by the downward displacement of water, four test-tubes and one gas bottle full of gas. Collect one full bottle of the gas by the downward displacement of air.

3. Stand the test-tubes and gas bottles mouth down on the table until ready to use. Fill the generator with water to stop the production of hydrogen.

Why is it important to have the lower end of the thistle-tube covered with acid?

What two functions are served by the thistle-tube?

What precaution must be taken if the delivery tube is made of rubber tubing or contains rubber connections?

PROPERTIES OF HYDROGEN

1. To test the solubility of hydrogen in water, place one of the test-tubes of gas mouth downward in a beaker of water and allow it to stand until the end of the period. In determining the solubility, lift the test-tube so that the level of any water inside the tube is the same as that in the beaker.

Why is it necessary to adjust the water levels in determining the solubility of a gas in this way?

What does this experiment show about the solubility of hydrogen in water?

2. Look at the hydrogen generator to make sure that it has ceased producing hydrogen and then you may light a burner. Bring a blazing splint to the mouth of one of the remaining test-tubes of gas held in an inverted position, and carefully observe the manner in which the gas burns.

What is characteristic about the way in which hydrogen burns?

Can hydrogen be identified by this characteristic behaviour in burning?

3. Lift the two remaining test-tubes from the table, quickly invert one of them, and hold the two tubes in the position shown in Fig. 32 for at least half a minute. At the end of this time bring a blazing splint to the mouth of each test-tube.

Describe the action in the test-tube held with the mouth up.

Describe the action in the test-tube held with the mouth down.

What does this experiment indicate regarding the relative densities of hydrogen and air?

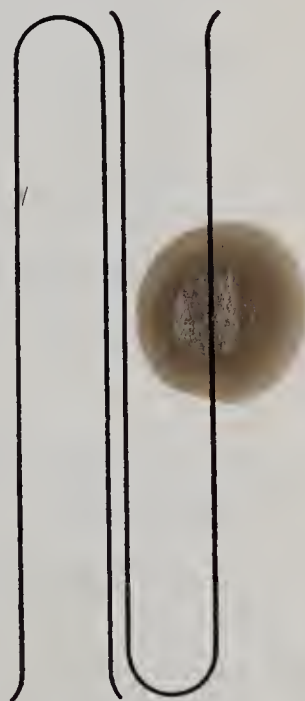


Fig. 32. Determining whether hydrogen is more or less dense than air.

4. Lift the bottle of hydrogen collected by the downward displacement of water from the table and keeping the mouth down (Fig. 33), slowly insert a long blazing wooden splint into the bottle. Leave the splint in the bottle for a short time and then slowly withdraw it from the bottle.

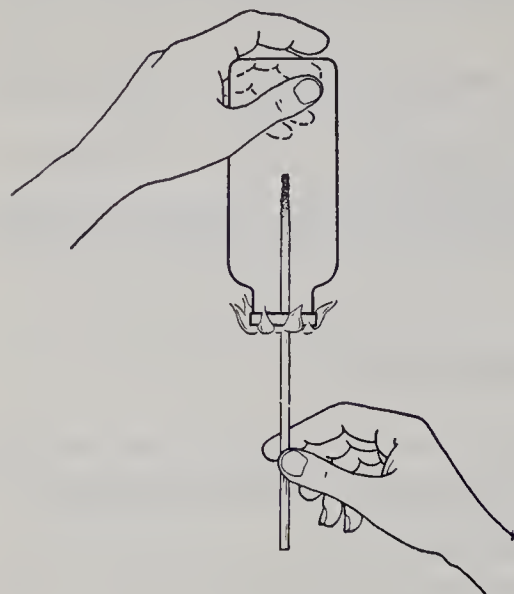


Fig. 33. Thrusting a blazing splint into a bottle of hydrogen.

Describe the action that takes place at the mouth of the bottle.

What happens to the burning splint? Does hydrogen support the combustion of the splint?

Name three important properties of hydrogen as determined by this experiment.

5. Note the appearance of the interior of the remaining bottle of hydrogen collected by the downward displacement of air. Place the bottle mouth upwards on the table and immediately bring a blazing splint to the mouth of the bottle. Observe the nature of the flame and any deposit on the inside of the bottle after the hydrogen has burned.

What is the reason for placing the mouth of the bottle uppermost in this experiment?

Describe the flame, and the appearance of the interior of the bottle after the gas has burned.

Suggest a name for the substance appearing inside the bottle.

Suggest a method for obtaining a sufficient amount of this substance to make a positive identification.

Why was this particular bottle collected by the downward displacement of air?

QUESTIONS

1. Write the word equation for the laboratory preparation of hydrogen.
2. List five properties of hydrogen.

✓ EXPERIMENT 27. (Demonstration) The burning of dry hydrogen at a jet.

1. Set up the apparatus as shown in Fig. 34, making sure that there is a plentiful supply of zinc in the flask and an additional

supply of acid available in the event that the supply of hydrogen dwindles. NOTE: The hydrogen generator shown in the diagram may be replaced by a Kipp generator (TEXT p. 47), which will supply a steady stream of hydrogen for a longer time than will the flask type generator.

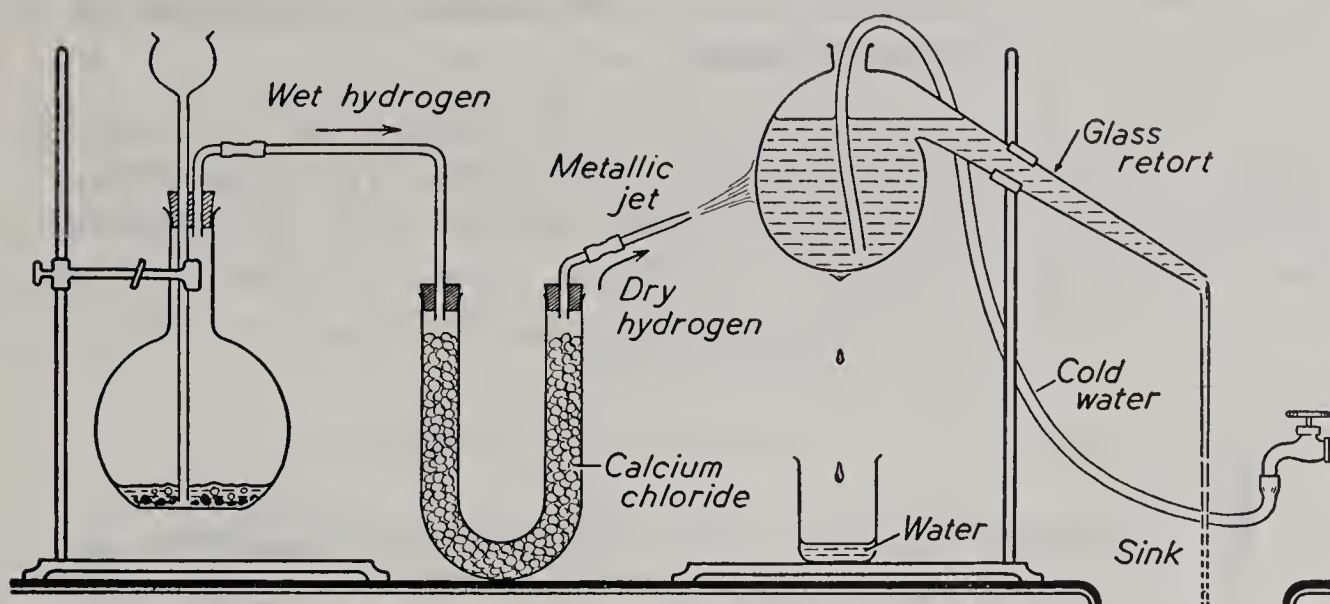


Fig. 34. Burning dry hydrogen at a jet.

2. Let the reaction continue for a short time so that the escaping hydrogen will displace the air originally present in the system. To be certain that pure hydrogen is issuing from the jet, collect a test-tube full of the gas and keeping the mouth down, bring it to the flame of a burner placed at some distance from the generating apparatus. When the gas burns quietly, with no explosion, bring the burning hydrogen to the jet and ignite the hydrogen at the jet with this flame. This is the only safe method of lighting the jet.

Explain the importance of this method of lighting the jet.

3. Let the flame of the burning hydrogen contact the surface of the glass retort and closely observe the surface. Test the substance formed with anhydrous copper sulphate.

Describe the flame of the burning hydrogen.

Describe the appearance of the outer surface of the retort.

What does the anhydrous copper sulphate test indicate?

Account for the formation of the deposit on the flask.

QUESTIONS

1. Why is it important to dry the hydrogen in Experiment 27? Does the calcium chloride in the tube show any evidence of being effective as a drying agent?

2. How does this experiment illustrate the “qualitative synthesis” of water?

3. Write a word equation to represent the burning of hydrogen in air.

✓ EXPERIMENT 28. (*Demonstration*) **The action of hydrogen as a reducing agent.**

1. Arrange the apparatus as in Fig. 35. Spread the cupric oxide along the tube evenly and make certain that the glass delivery tube does not become clogged when placing it in the test-tube. Let the reaction continue a sufficient length of time to drive the air out of the apparatus and then bring the flame under the part

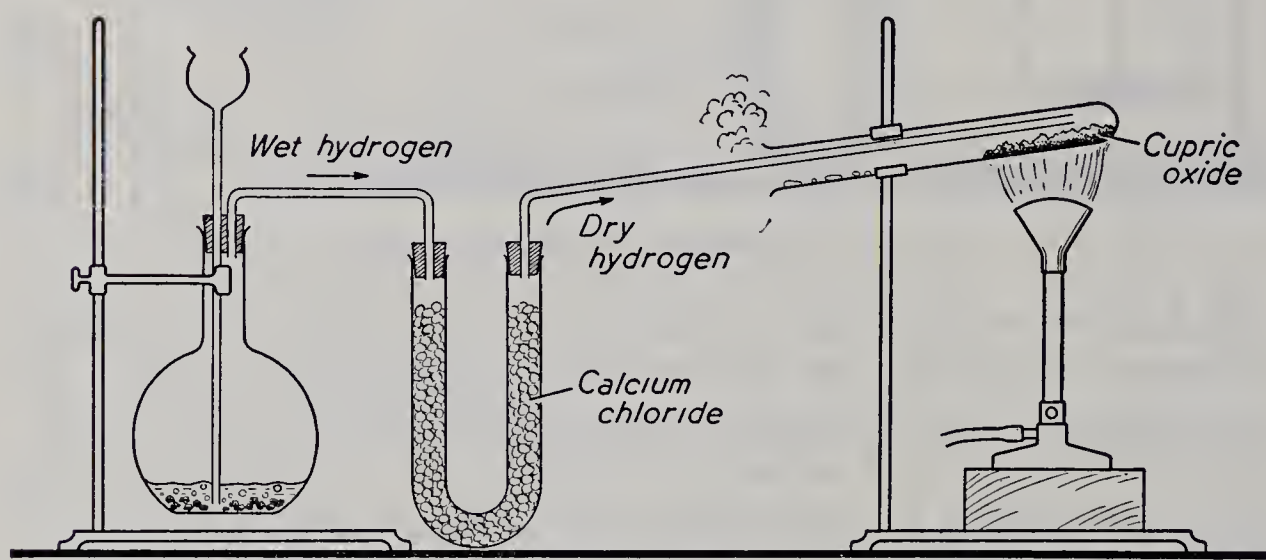


Fig. 35. Reducing cupric oxide with dry hydrogen.

of the tube containing the cupric oxide. Do not bring the burner flame near the open end of the test-tube. Continue heating until the action within the test-tube ceases. Then remove the burner and carefully withdraw the glass tube making sure that it does not touch the sides of the test-tube.

Give two reasons for keeping the burner flame away from the mouth of the test-tube.

What changes take place in the cupric oxide?

What is the nature of the deposit near the mouth of the test-tube?

List any other observations noted.

2. Allow several drops of the liquid formed in the test-tube to fall on some anhydrous copper sulphate. When the test-tube has cooled sufficiently, tap the material in the tube onto a watch-glass and after observing the material closely, add a few drops of concentrated

nitric acid. Rinse out the test-tube with water, examine any deposit on the inside of the tube and add a few drops of concentrated nitric acid.

What is the liquid formed in the cool part of the test-tube? Account for its formation.

What is the appearance of the mixture on the watch-glass?

What changes take place when the concentrated nitric acid is added?

What does this test indicate? (TEXT p. 252)

Name three possible substances present in the mixture.

What is the deposit adhering to the inside of the test-tube? Account for its formation.

QUESTIONS

1. Write two possible word equations for the action of hydrogen on cupric oxide.
2. Explain the action of hydrogen as a reducing agent in this experiment.

UNIT 5

Water

EXPERIMENT 29. The purification of water.

1. Add a tablespoonful of garden soil to 500 ml. of water in a beaker, and stir. Pour half of the mixture into another beaker so that you now have two beakers containing muddy water. Allow one beaker to stand for 20 minutes.

2. Add 10 ml. of a 10% solution of aluminum sulphate to the second beaker containing the muddy water and stir. Then add 10 ml. of a saturated solution of calcium hydroxide (limewater), stir and allow to stand.

Compare the rates of the settling (sedimentation) of the soil particles in Procedures 1 and 2. Explain the difference in rate.

3. Stir the water in the first beaker again and filter a small quantity of it. Compare the clarity of the filtrate with that of the water above the sediment in the untreated beaker.

Which is clearer: filtered water, water from which the sediment has settled by gravity, or water which has been clarified by chemical action?

QUESTIONS

1. What effect has the addition of certain chemicals on the speed of settling or sedimentation? Explain how this effect is produced.

2. Of the above three methods, which is the most efficient from the standpoint of the clarity of the water produced? Which is most efficient from the standpoint of time required?

3. Is the water produced either by sedimentation, or by filtration actually pure water? Explain.

4. Would any of the above methods make water safe for drinking? Explain.

EXPERIMENT 30. (*Demonstration*) The distillation of water.

1. Into a beaker containing 50 ml. of water put a few crystals of potassium permanganate and stir. Add a teaspoonful of soil and stir again.

What happens to the potassium permanganate? Was the water in the beaker clear after the potassium permanganate was added?

Was the water clear after the soil was added?

2. Set up a distilling apparatus as shown in Fig. 36. Pour the mixture from Procedure 1 into the flask (A), connect the lower rubber

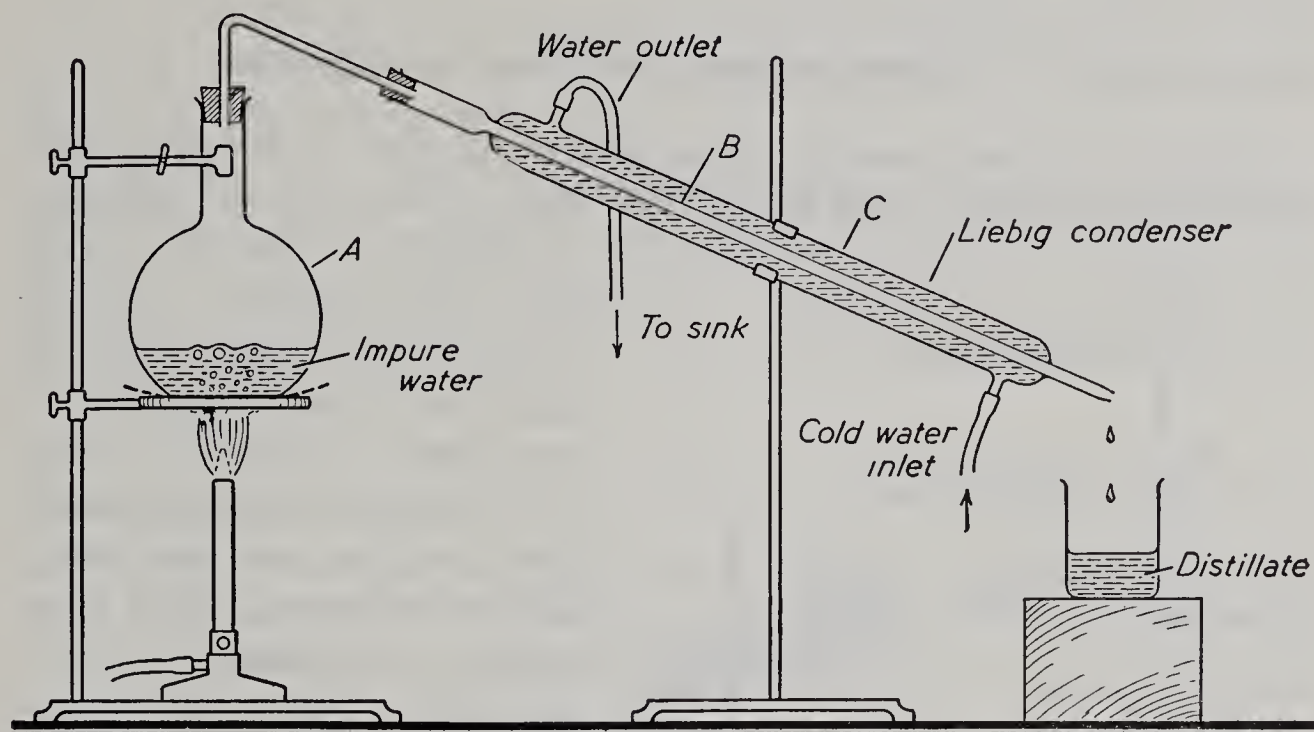


Fig. 36. The distillation of water using a Liebig condenser.

tubing leading to the condenser to a water tap, and lead the other tubing to the sink. Turn on the tap so that a steady stream of cold water flows through the glass jacket (C) of the condenser. Heat the mixture in the flask until it boils, and continue to heat gently for ten minutes. Collect some of the distillate in a beaker as shown. Examine and taste the distillate. Evaporate about 2 ml. of the distillate to dryness in either an evaporating dish or Pyrex test-tube.

Compare the properties of the mixture in the flask with the properties of the distillate.

What is observed when some of the distillate is evaporated to dryness? What does this show about the effect of distillation on (a) dissolved, and (b) suspended substances?

Name the two changes in state involved in the distillation process.

Is the distillate pure water? Explain.

Why is the cold tap water allowed to enter at the lower end of the condenser?

QUESTIONS

1. Compare water obtained by the process of distillation with water obtained by filtration as to (a) clarity, and (b) purity.
2. Why is distillation often called a "double" process?
3. Why is the process of distillation not used in obtaining a pure water supply for a city?

EXPERIMENT 31. (*Demonstration*) **Fractional distillation.**

Replace the flask used in Experiment 30 with a distilling flask which has a side-arm, and fit the flask with a thermometer as shown in Fig. 37. Pour a mixture of equal volumes of ethyl alcohol and water into the flask and then connect the condenser as shown. Heat the flask until the liquid boils and continue to heat slowly for ten minutes. Take temperature readings every two minutes. Collect five separate samples of the distillate and smell each sample.

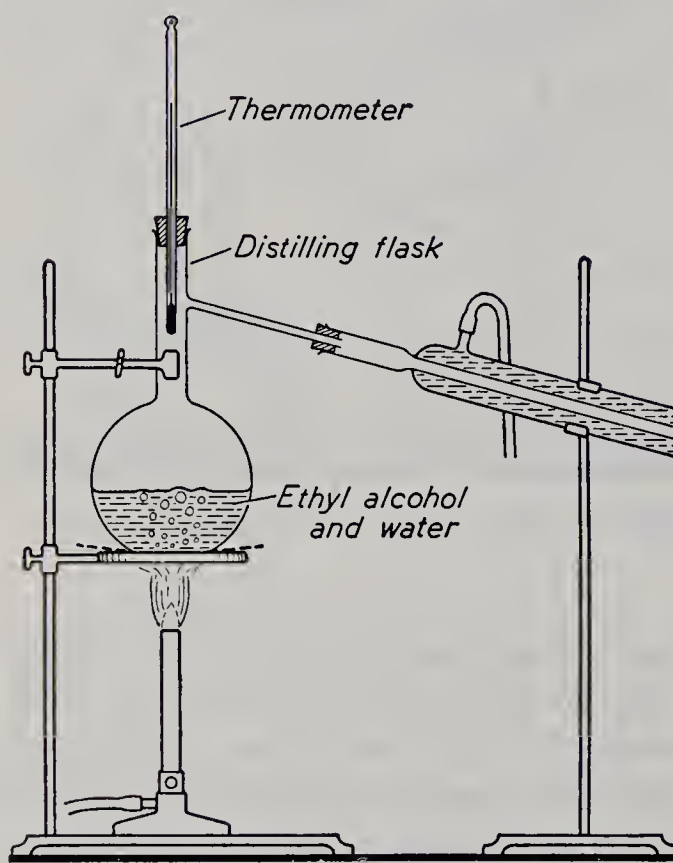


Fig. 37. The fractional distillation of a mixture of ethyl alcohol and water.

Pour a mixture of equal volumes of ethyl alcohol and water into the flask and then connect the condenser as shown. Heat the flask until the liquid boils and continue to heat slowly for ten minutes. Take temperature readings every two minutes. Collect five separate samples of the distillate and smell each sample.

At what temperature does the initial boiling occur? Is there any change in the thermometer reading as the boiling continues?

Which of the two liquids in the distilling flask boiled at this temperature?

Which distillate contains the highest percentage of alcohol? (TEXT p. 60)

What name is given to the process whereby two liquids of different boiling points are separated?

How could you prove that each sample of distillate differs considerably from the liquid originally in the distilling flask?

QUESTIONS

1. Explain the use of the word "fractional" in referring to a distillation of this type.

2. Briefly explain how a mixture of liquids, such as are found in crude oil, may be separated by fractional distillation. (TEXT p. 187)

✓ EXPERIMENT 32. The properties of water.

1. Recall experiments performed in earlier grades and record each of the following properties of water in your notebook: colour, odour, taste, conductivity of heat and of electricity, freezing point (760 mm. pressure), boiling point (760 mm. pressure), heat of vaporization, heat of fusion, and density at 4°C.

2. Review your observations for the experiments completed in Units 2 and 4. In studying these units, emphasis was placed on the chemistry of oxygen and hydrogen. In this review, concentrate on the properties of water and list all the chemical properties illustrated by these experiments.

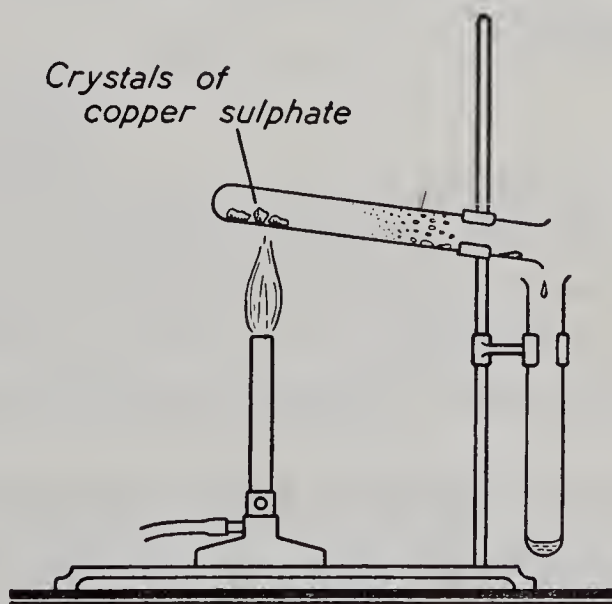
✓ 3. Add a pinch of anhydrous copper sulphate to a few drops of water on a watch-glass.

Describe the appearance of anhydrous copper sulphate, (a) before water is added and (b) after water is added.

When water is added to anhydrous copper sulphate, is there a physical or chemical change? Explain.

✓ EXPERIMENT 33. A study of hydrates.

1. Gently heat a few crystals of copper sulphate (bluestone) in a clean dry test-tube clamped with the mouth slightly lower than the rest of the tube. Clamp a second test-tube vertically as shown in Fig. 38. Increase the temperature gradually until the crystals are being strongly heated. Test a drop of the material caught in the vertical tube with anhydrous copper sulphate.



What two changes are observed in the copper sulphate crystals? Fig. 38. The dehydration of bluestone.

Describe the substance found in the vertical test-tube and on the cooler portion of the sides of the heated test-tube. Account for its formation. What is this liquid? How did you identify it?

Does the heating of a hydrate represent a physical or a chemical change? Explain.

2. Gradually move the Bunsen burner forward towards the mouth of the upper test-tube and continue heating until the tube and its contents are dry. Cool the test-tube. Shake out the contents of the tube on a piece of paper and examine the substance. Add a drop of water to this substance.

Describe the appearance of the substance on the paper.

What is observed when water is added to the substance formed by heating bluestone? What is this substance?

Write a word equation to represent the reaction that takes place when bluestone is strongly heated. What name is given to this type of reaction?
(TEXT p. 62)

Write a word equation to represent the reaction that takes place when water is added to anhydrous copper sulphate. What name is given to this type of reaction?

3. Repeat Procedures 1 and 2 using a few crystals of each of the following substances in separate dry test-tubes: alum, sodium chloride, and sodium carbonate.

Describe what happens in each case.

Which of these substances decomposes when heated, to form an anhydrous salt and water? What are such substances called?

Would this experiment indicate that all crystalline substances are hydrates?

QUESTIONS

1. Define a hydrate and name five hydrates.

2. Why is "water of hydration" more suitable to express the water present in a hydrate than is "water of crystallization"?

3. Suggest an experiment that could be performed to find the percentage of water of hydration in bluestone.

✓ EXPERIMENT 34. Efflorescence and deliquescence.

1. Weigh about a teaspoonful of sodium thiosulphate crystals (hypo) on a watch-glass and allow the watch-glass and contents to stand for most of the laboratory period. Then weigh again.

Describe any change in the appearance of the crystals.

Account for any change in weight observed.

2. Weigh about a teaspoonful of granular anhydrous calcium

chloride on a watch-glass and allow the watch-glass and contents to stand for most of the laboratory period. Then weigh again. (While waiting continue with Procedure 3.)

Did the calcium chloride change in appearance? If so account for the change.

Did the weight of the calcium chloride change? If so why?

3. (For this procedure the class may be divided into groups.) On separate watch-glasses place about one-quarter of a teaspoonful of each of the following nine substances; "hypo" crystals, granules of anhydrous calcium chloride, pellets or lumps of sodium hydroxide, crystals of sodium nitrate, crystals of potassium nitrate, crystals of magnesium chloride, washing soda crystals, crystals of Glauber's salt, and anhydrous copper sulphate. Place the watch-glasses and their contents in a convenient place until the next laboratory period and examine the substances again.

Describe any changes that occur in each of the substances used.

Compare the "hypo" after exposure to air for a day with the "hypo" exposed for 20 minutes as in Procedure 1.

What other substances behaved as did "hypo"? What are such substances called? (TEXT p. 62)

Compare the calcium chloride after exposure to air for a day with the calcium chloride exposed for 20 minutes as in Procedure 2.

What other substances behaved as did calcium chloride? What are such substances called? (TEXT p. 63)

How do you explain the behaviour of the anhydrous copper sulphate?

QUESTIONS

1. Distinguish between efflorescence and deliquescence.¹ Are both of these phenomena types of chemical reactions? Explain.
2. Explain why deliquescent substances make good drying agents.
3. Distinguish between dehydration and efflorescence.

EXPERIMENT 35. (*Demonstration*) The electrolysis of water.

1. A Hoffman apparatus is used (Fig. 39). To about 250 ml. of water add about 5 ml. of concentrated sulphuric acid (TECH. 3). With both stop-cocks open, slowly pour the acidulated water into the reservoir until the two outer tubes are filled up to the taps. Avoid getting any liquid in the jet tubes above the taps. Close the taps and add enough acidulated water to half-fill the reservoir. Paste a

paper label on the reservoir so that its upper edge marks the surface level of the liquid it contains. Connect the platinum electrodes to a source of *direct* current. If 3 dry cells in series are used the electroly-

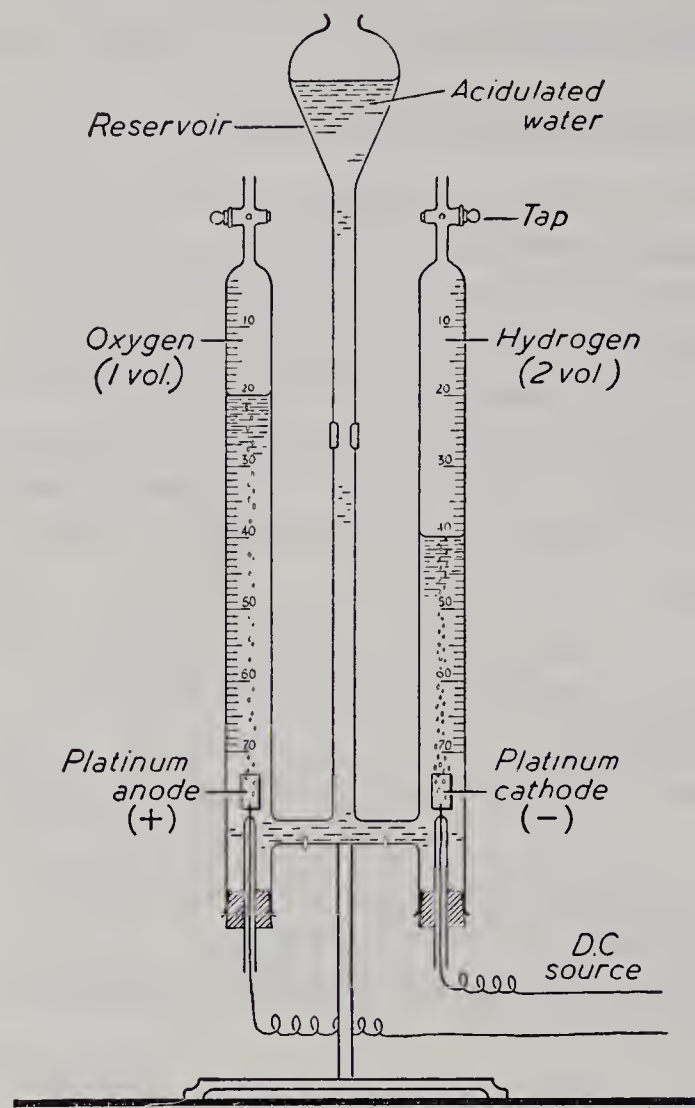


Fig. 39. The electrolysis of water using a Hoffman apparatus.

sis should be started before the class period to insure the accumulation of enough gases for testing. If a D.C. source of higher voltage (e.g. 110 volts) is used, a lamp or some other resistance must be connected in series in the circuit. The circuit which includes a switch and the Hoffman apparatus should be carefully examined and the direction of the current determined before beginning the experiment.

Close the switch and allow the current to pass through the acidulated water until the tube above the cathode contains about 20 ml. of gas. Open the switch and then accurately read the volume of gas collected (a) above the cathode, and (b) above the anode. Observe the new level of the liquid in

the reservoir. Hold an inverted test-tube over the jet that is above the cathode and open the tap. Keep the tap open until all the gas escapes into the test-tube but do not permit any liquid to escape at the jet. Bring a lighted splint to the mouth of the inverted test-tube. Test the gas above the anode by holding a glowing splint above the jet at the same time opening the tap. Do not allow any liquid to escape at the jet. Observe the level of the liquid in the reservoir after the tubes are again filled with liquid.

Describe what was observed when the switch was closed.

What was the volume of the gas collected (a) above the cathode, (b) above the anode?

Name the gas (a) above the cathode, (b) above the anode. How were these gases identified? What were the relative volumes of the two gases?

From a comparison of the liquid levels in the reservoir before the electrolysis began, and after the tubes were again filled with liquid, what conclusion may be made about the amount of water that has been electrolysed?

2. If time permits repeat the above experiment.

QUESTIONS

1. Make a diagram of a Hoffman apparatus connected in a circuit as used for the electrolysis of water. Label all the essential parts.
2. Why is sulphuric acid added to the water?
3. Explain the changes in level of the acidulated water in the reservoir as the electrolysis takes place and as the gases are allowed to escape from the jets.
4. What kind of energy was used to decompose water into the two gaseous elements?

EXPERIMENT 36. (*Demonstration*) The synthesis of water (by volume).

1. A eudiometer tube is used (Fig. 40). Because of the expense of the large volume of mercury required to fill the jar and tube, and the greater danger of breakage, it is suggested that water be used in place of mercury. Examine a eudiometer tube so that you will be able to describe it in detail. Fill the tube with water and invert it in a tall hydrometer jar also filled with water. Pass a small volume (e.g. 15 ml.) of hydrogen (from a hydrogen generator) into the tube. Lower the tube until the level of the water is the same inside and outside the tube. Accurately read the volume of hydrogen and record it. Add about the same volume (15 ml.) of oxygen (from an oxygen cylinder or generator) and again adjust the tube until the water levels are the same inside and outside the tube. Read and record the volume of the mixed gases.

CAUTION: Make sure that the two gases are present in volumes that assure some residual gas in the eudiometer after the explosion. Because the eudiometer may shatter, no one except the instructor should be within 10 feet of the apparatus.

Connect the platinum terminals to an electric circuit containing three or four dry cells connected in series, an induction coil, and a switch (open). Grasp the eudiometer with a towel and press it

firmly on a rubber pad placed in the bottom of the hydrometer jar, and at the same time press the hand on the top of the jar. Close the

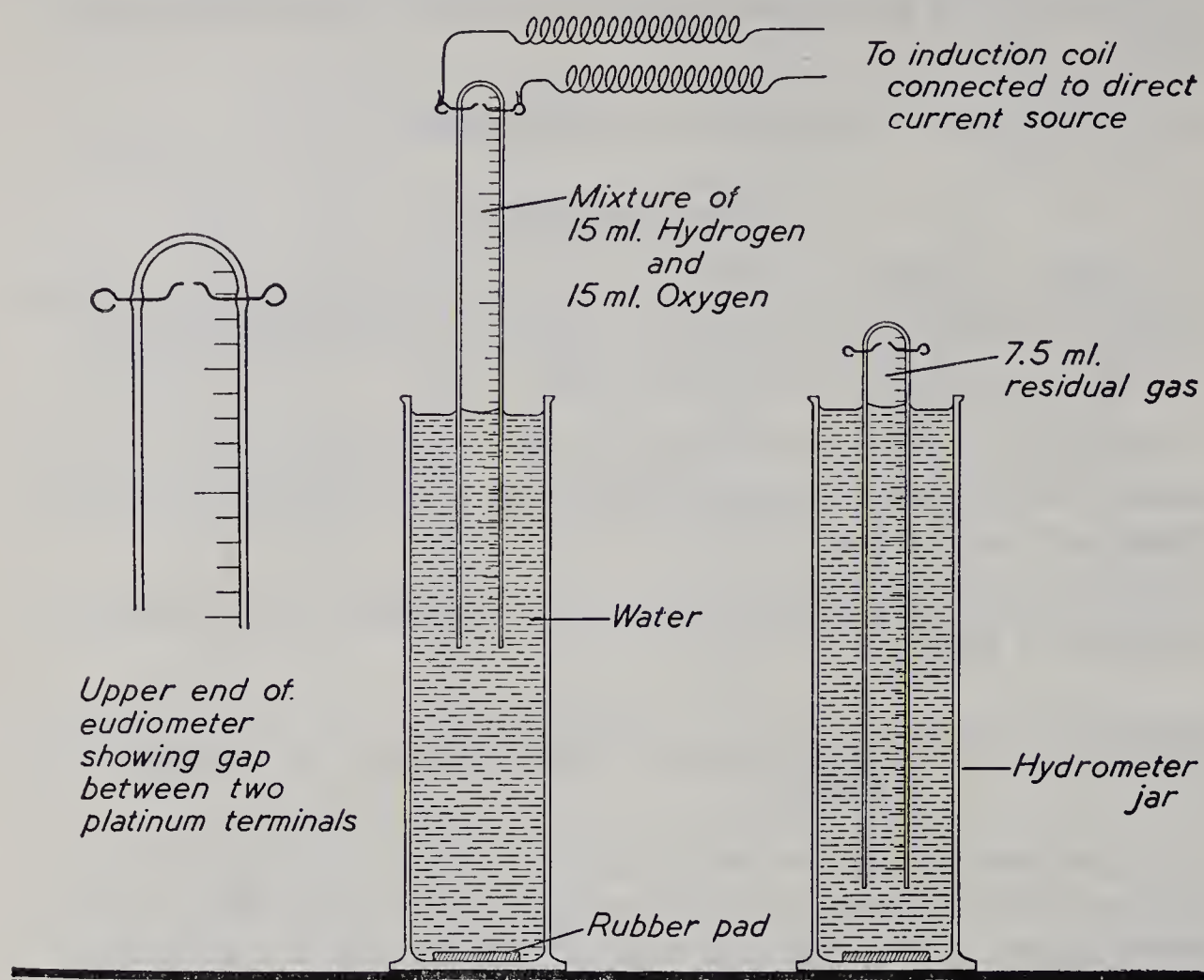


Fig. 40. Synthesis of water by volume in a eudiometer.

switch. Observe any changes that occur inside the tube just at the instant the switch is closed and immediately after. Wait until the tube cools and then adjust the levels as before and measure the volume of the residual gas. Pass a spark through the residual gas.

What is observed at the moment the circuit is closed? Explain the downward movement of the water.

What causes the sudden uprush of water that followed shortly after the explosion?

Why does no explosion occur when a spark is passed through the residual gas?

2. Place the hand or a small glass plate under the mouth of the eudiometer tube, which now contains only the residual gas and water, and remove the tube and contents from the hydrometer jar. Place the tube in an upright position, remove the hand and immediately insert a burning splint.

What is the residual gas?

What is the ratio by volume of the hydrogen and oxygen which disappeared? What happened to these gases? What was observed during the experiment to support your conclusion?

3. Repeat the experiment but this time use approximately 25 ml. of hydrogen and 7 ml. of oxygen. Tabulate your observations as follows:

1. Volume of hydrogen used ml.
2. Volume of the mixture of hydrogen and oxygen _____ ml.
3. Volume of oxygen used (2—1) ml.
4. Volume of residual gas ml.

4. After the explosion has taken place and the volume of the residual gas has been measured accurately, add about 5 ml. of hydrogen to the eudiometer tube. Pass a spark through the gas.

Explain why no explosion occurs.

Suggest how you might verify your identification of the residual gas.

QUESTIONS

1. What would be one advantage of using mercury to fill the tube and jar in preference to water?

2. Was the reaction between oxygen and hydrogen endothermic or exothermic? Give reasons for your answer.

3. Why should care be taken not to have the exact proportions in which hydrogen and oxygen unite, present in the eudiometer tube? (TEXT p. 66)

EXPERIMENT 37. (*Demonstration*) **The synthesis of water (by weight). (Dumas' experiment)**

The apparatus used (Fig. 41) consists of three drying tubes containing granular calcium chloride, a combustion tube and porcelain boat, a Bunsen burner fitted with a flame spreader, and a hydrogen generator (not shown in the diagram). Weigh together the drying tubes (C) and (D) and record the weight. Nearly fill the boat with cupric oxide, place it in the combustion tube and weigh the tube and contents together. Record this weight. Assemble the complete apparatus and connect it to a hydrogen generator. Pass hydrogen through the assembled tubes until all the air has been driven from them. To make certain that hydrogen has replaced the air in the system, test at the exit tube for pure hydrogen (see Experiment 27). When hydrogen only is escaping, then, *and not until then*,

heat the combustion tube as shown in Fig. 41. Heat strongly for 5 minutes and then remove the flame and allow the tube and

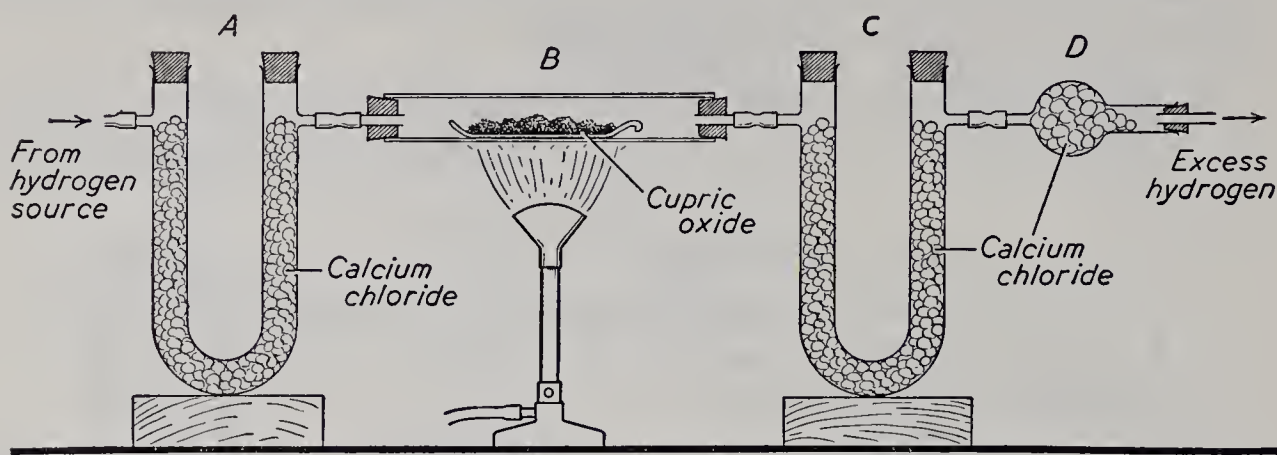


Fig. 41. Finding the percentage composition of water by weight.

contents to cool. Remove (C) and (D) together and weigh them again. Remove the combustion tube with its contents and weigh again. Record the observations as follows:

1. Weight of (B) before reduction.....gm.
2. Weight of (B) after reduction.....gm.
3. Weight of oxygen used (1—2).....gm.
4. Weight of tubes (C+D) before
experiment.....gm.
5. Weight of tubes (C+D) after
experiment.....gm.
6. Weight of water formed (5—4)gm.
7. Weight of hydrogen in the water
formed (6—3).....gm.

From these observations calculate the percentage composition of water by weight.

Explain why (1 — 2) gives the weight of the oxygen used in forming the water obtained.

Explain why (5 — 4) gives the weight of the water that is formed by the reduction.

QUESTIONS

1. Why is the drying tube (A) used? Why was it not weighed before and after the reduction?
2. Why is the heating of the combustion tube delayed until the tubes are free of air?
3. What is the percentage of error in your experiment? Compare your result with that of Morley. (TEXT p. 68)

UNIT 6

Solutions

✓ EXPERIMENT 38. The formation of solutions.

1. Add a small crystal of potassium permanganate to 10 ml. of water in a test-tube and shake until the crystal disappears. Note the appearance of the liquid. Add a few more small crystals and shake again until the crystals disappear.

What is the colour of a mixture of potassium permanganate and water?

Why is this mixture called a solution?

Is the solution clear or turbid?

What effect has the addition of more crystals on the colour of the solution?

Is there any effect on the clarity of the solution?

2. Place 50 ml. of water in a graduated cylinder. Weigh on a piece of paper 6 gm. of ammonium chloride. Take and record the temperature of the ammonium chloride and then of the water. Add the salt to the water and shake until solution is effected, making sure that none of the solution escapes. Take the temperature of the solution and measure its volume.

What is the colour of the solution of ammonium chloride?

Is there a change in temperature? If so, state the change.

Is there any change in the volume of the liquid?

3. (*Demonstration*) Place about 75 ml. of water in a gas-measuring tube (capacity 200 ml.). Slant the tube at an angle and slowly add about 75 ml. of alcohol in such a way that very little mixing of the two liquids occurs. Hold the tube upright and record the volume accurately. Put a stopper in the tube and shake for a few minutes. Record the volume.

What evidence is there that a solution has been formed?

What was the total volume of the liquids before mixing? What was the volume after mixing?

Does the solution occupy as great a volume as the two components occupy separately?

4. Allow water to run from the "cold-water" tap until the water is at the coldest possible temperature. Fill a beaker with cold water and set it aside in a warm place for several hours. Examine the inside walls of the beaker.

Account for the bubbles that are seen clinging to the walls of the beaker. Of what are these bubbles composed?

What does this experiment show about the relative solubilities of air in cold water and in warmer water? Is this property characteristic of all gases? (TEXT p. 73)

✓ EXPERIMENT 39. **Factors affecting the rate of solution.**

1. Half fill two test-tubes with water. To each add 1 gm. of copper sulphate crystals and allow one tube to stand undisturbed. Cover the mouth of the second tube with your thumb and shake vigorously.

Is copper sulphate soluble in water? Give reasons for your conclusion.

What effect has agitation on the rate with which the distribution of the solute throughout the solvent takes place? Explain.

Do you think that the copper sulphate in the undisturbed test-tube would eventually dissolve? Explain.

2. Half fill two test-tubes with water. To one add 2 gm. of copper sulphate crystals and to the other the same quantity of powdered copper sulphate made by grinding the crystals with a mortar and pestle. Successively shake each tube and note the time required to dissolve the copper sulphate in each.

Which dissolves more rapidly? Why is there a difference in the rate of solution?

What method of speeding up the rate of solution is illustrated by this experiment?

3. Half fill two test-tubes with water. Add 2 gm. of copper sulphate crystals to each. Let one stand undisturbed and heat the other in a Bunsen flame.

What effect has an increase in temperature on the rate of solution?

✓ EXPERIMENT 40. **Factors affecting solubility.**

1. Half fill six test-tubes with water and place them in a large beaker or a test-tube rack. To each test-tube add one of the following: 1 gm. sugar, 1 gm. sodium chloride, 1 gm. calcium carbonate, 2 ml. alcohol, 2 ml. glycerine, and 2 ml. kerosene. Shake each tube and observe the results.

Which of these substances are insoluble? Which are immiscible? Which are soluble in water?

Which of the following factors is responsible for the difference in solubility: the nature of the solvent, the nature of the solute, or the conditions of the experiment?

2. Half fill three test-tubes with water, alcohol, and carbon tetrachloride respectively. Add a small crystal of iodine to each, shake and observe the results.

Does the iodine dissolve in all three liquids? Explain.

Which of the following factors is responsible for the difference in solubility: the nature of the solvent, the nature of the solute, or the conditions of the experiment?

3. Repeat Procedure 2 using small lumps of margarine in place of iodine crystals.

Does the margarine dissolve in all three liquids? Explain.

Which factor is responsible for the difference in solubility: the nature of the solvent, the nature of the solute, or the conditions of the experiment?

Which of the three liquids used would be most satisfactory for removing grease stains from cloth? Why?

4. Weigh 10 gm. of ammonium chloride on a piece of paper and then divide the salt into five portions of approximately 2 gm. each. Put 10 ml. of water into a large test-tube, add a 2-gm. portion of the salt, and shake the test-tube until the salt dissolves. Repeat this process until a small amount of the solute remains undissolved in the test-tube. Heat the test-tube until the solution is warm (about 50°C.) and note the effect on the excess salt. Add another portion of the salt and shake again. Heat again until the solution boils. Add the remaining portions of the salt, shaking after each addition. Cool to room temperature.

Which factor is responsible for the difference in solubility: the nature of the solvent, the nature of the solute, or the conditions of the experiment?

How does temperature affect the solubility of ammonium chloride?

Which of the solutions examined were unsaturated? Which were saturated? Explain.

What occurred when the hot concentrated solution was cooled? Explain.

QUESTIONS

1. Distinguish between "rate of solution" and "solubility".

2. A salt is shaken in water for some time and when agitation

ceases some of the salt settles. What is the condition of the solution above the salt? How may the excess salt be put into solution?

3. What four factors are always implied when the term *solubility* is used? Define solubility. (TEXT p. 75)

4. Suggest how you would determine the solubility of ammonium chloride at 60°C.

5. Distinguish between the terms "clear" and "colourless". Are all solutions clear? Are all solutions colourless?

✓ **EXPERIMENT 41. The determination of the solubility of a salt.**

It is suggested for this experiment that the solubility of several salts be determined by distributing various salts to different groups. The method given here for common salt (sodium chloride) would apply, except for quantities stated, to each of the salts selected.

1. Shake together in a *large* test-tube 30 ml. of water and 15 gm. of sodium chloride until no more salt dissolves. Take the temperature of the saturated solution.

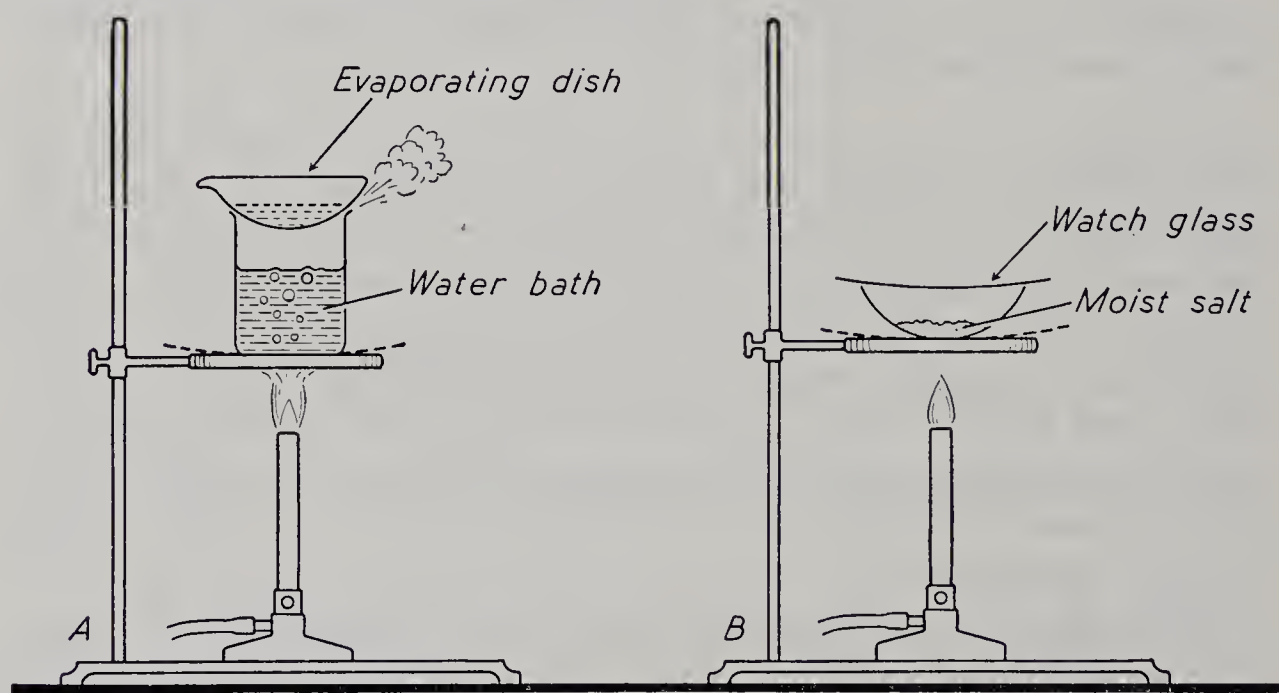


Fig. 42. Evaporating a solution. A, first stage, using a water bath. B, second stage, evaporating to dryness.

2. Weigh together a clean, dry evaporating dish and a watch-glass cover. Record this weight. Filter about 20 ml. of the saturated solution into the evaporating dish, place the cover on it and weigh. Record this weight. Remove the cover and place the evaporating dish over a beaker of boiling water (or on a sand-bath) to evaporate the solution (Fig. 42A). Evaporate until nearly

dry and then remove the evaporating dish. Wipe the water from the outside of the dish and then place it on a wire gauze supported on a ring stand.

3. Again cover the dish with the watch-glass and heat the dish carefully with a small Bunsen flame (Fig. 42B) until thoroughly dry. Be sure that all traces of moisture are driven not only from the evaporating dish, but from the under surface of the cover as well. Allow the dish to cool to room temperature and then weigh together the dish, contents and cover, and record the weight. Record your data as follows:

1. Temperature of the saturated solution _____°C.
2. Weight of evaporating dish and cover _____gm.
3. Weight of evaporating dish and cover
and solution _____gm.
4. Weight of saturated solution (3-2) _____gm.
5. Weight of evaporating dish and
cover and residue _____gm.
6. Weight of sodium chloride dissolved (5-2) . _____gm.
7. Weight of water in the saturated
solution (4-6) _____gm.

Calculate the weight of sodium chloride required to saturate 100 gm. of water at the temperature of the solution in your experiment.

How does your result compare with that shown by the solubility curve on page 77 of the text?

What is your percentage of error? (See page 3.)

Define "the solubility of a salt".

QUESTIONS

1. Why was the evaporating dish covered with a watch-glass during the last stage of evaporation?
2. What sources of error may have caused your result to be inaccurate?
3. Suggest how you could determine the solubilities of a certain salt at different temperatures.

EXPERIMENT 42. To make a supersaturated solution.

1. Put crystals of sodium thiosulphate (hypo) to a depth of four inches in a clean dry test-tube taking care that no crystals stick to the sides of the tube (TECH. 8). Add 10 (no more) drops of water

and gently shake the tube for a few minutes. Note any change in temperature and examine the contents. Gently warm the tube long enough to get all the "hypo" into solution. Place some absorbent cotton in the mouth of the test-tube and hold the tube in a stream of cold water from the tap until it is cooled to approximately its original temperature. Examine the contents of the tube.

What evidence have you that "hypo" is soluble in cold water? Does it appear to be very soluble? Is there any change in temperature when solution occurs?

What effect has heating on the solubility of "hypo" in water?

Why is the cold solution you have just made called a supersaturated solution?

At what stage during this procedure is the solution saturated? Give a reason for your answer.

Is the hot solution saturated before cooling? How could you have tested it to find out?

2. Hold the test-tube that contains the supersaturated solution so that it may be viewed against a light background, and add first a crystal of blue vitriol and one minute later a small crystal of "hypo" to the solution and observe. Note any temperature change.

What was the effect of adding the copper sulphate crystal?

What happened when the "hypo" crystal was added? Was there any change in temperature? What kind of solution remains?

From knowledge gained in previous experiments, why would you expect to have an increase in temperature when crystallization occurs?

QUESTIONS

1. Describe, using diagrams to illustrate your answer, how you would make a supersaturated solution of Glauber's salt. What would happen if a small crystal of sodium sulphate were added to the solution?

2. When a crystal of the same solute is added to a supersaturated solution what is the condition of the solution that results? Explain.

3. In preparing the supersaturated solution of "hypo", why was absorbent cotton used as a plug in the end of the test-tube?

4. Account for the presence of such a large volume of solution when only a few drops of water were used in the preparation of the supersaturated solution.

EXPERIMENT 43. (*Demonstration*) To make a large crystal of alum.

1. Dissolve about 30 gm. of powdered potassium aluminum sulphate in 100 ml. of warm water (about 50°C.) in a beaker. Allow the solution to stand over night and then filter it into another clean beaker. Select the most perfect crystal that has formed during the night and place it in the clear filtrate. Cover the beaker with a filter paper and again allow the solution to stand over night. On the following day remove the crystal and clean it by scraping off any small nodules that may be adhering to it, and then by wiping the crystal with a towel. Filter the solution to remove any new crystals that may have formed, and then replace the cleaned crystal in the filtered solution.

2. The above procedure should be repeated every day for about two weeks. The crystal should be turned from one crystal face to another at intervals so that the growth of the crystal will be uniform.

After the final cleaning, if the crystal is to be handled it may be covered with a coat of shellac.

UNIT 7

The Gas Laws

EXPERIMENT 44. (*Demonstration*) **The measurement of air pressure.**

1. Select a thick-walled glass tube about three feet in length, and closed at one end. Carefully fill this tube with mercury. (During the filling process, it is advisable to *tap* the tube gently at intervals to prevent the retention of trapped air bubbles.) When it is filled, hold a finger firmly over the open end of the tube; then invert the tube. Keep the finger in position until the end of the tube is submerged beneath the surface of mercury in a dish (cistern). Support the tube in a vertical position, and then remove the finger.

State what is observed when the finger is removed.

Why does the mercury "oscillate" for a time? Why does it eventually come to rest?

What is present in the space in the tube above the mercury?

2. Clamp a metre stick beside the tube so that the zero extends below the mercury in the dish. Read the metre stick (in both millimetres and inches) at the place where it enters the mercury and also directly opposite the surface of the mercury in the tube, and record these measurements. (When measuring a mercury column, always read the top of the meniscus.)

What is the height of the mercury column in (a) inches, (b) millimetres?

What does this height measure? Explain how "units of length" are used to measure pressure. (TEXT p. 90)

Why is such a tube called a Torricellian tube? Why is it also called a barometer?

3. Examine the barometer at different times every day for a period of one week and record the heights of the mercury column.

Explain the variations noted.

EXPERIMENT 45. (*Demonstration*) To determine the relation between the volume and the pressure of a gas. (Boyle's Law)

1. Use a glass tube at least 4 mm. in diameter and 96 cm. long. The tube is bent in a U-form so that one arm is much shorter than the other. The shorter arm is sealed. Such a tube is often called a Boyle's Law tube (Fig. 43). (If such a tube is not available, two straight tubes, one short and sealed at one end and the other longer and open at both ends, may be connected by a heavy rubber tube and substituted for the Boyle's Law tube.)

Support the tube vertically on a stand.

2. Pour mercury into the tube until it just fills the bend. Then manipulate the tube back and forth until the mercury reaches the same level in both arms. This is done so that the pressure of the enclosed air in the shorter arm of the tube will be at the same pressure as that of the atmosphere.

Clamp a metre stick to the stand so that it is in a vertical position, and so placed that the height of both arms of the tube can be measured. Record the position of the mercury level (in centimetres) in each arm of the tube, and the volume (represented by the height in centimetres) of the air enclosed in the shorter arm of the tube.

Add mercury to the tube until the mercury in the longer arm is about 15 cm. above the mercury in the shorter arm and record the height of the mercury levels in both arms of the U-tube, as indicated by the readings on the metre stick. Record also the volume (height) of the enclosed air.

Increase the height of the mercury column in the longer arm by

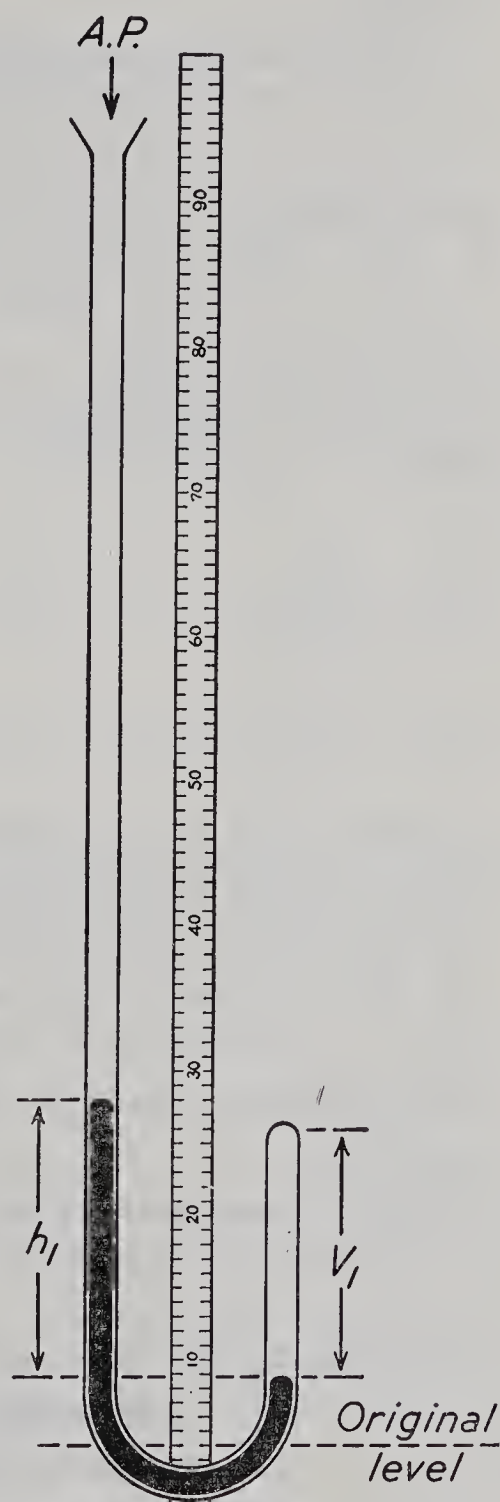


Fig. 43. Using a Boyle's Law tube.

about 15 cm., by adding more mercury to the tube. Observe and record the positions of the mercury levels in each arm of the tube. Record the volume of the enclosed air. Continue to add more mercury at intervals until the longer arm is nearly filled. Each time mercury is added, take a reading of the mercury levels in both the longer and shorter arms of the tube and record the volume of the enclosed air.

3. Read a mercury barometer (in cm.) and record the reading. Let this be represented by A.P. (atmospheric pressure).

4. Let V , V_1 , V_2 , V_3 , etc., represent the volumes of air enclosed in the shorter arm of the tube at successive observations as indicated by the lengths of the air column.

Let h_1 , h_2 , h_3 , etc., represent the *differences* in mercury levels in the two arms of the tube at successive observations.

Complete the following table:

Volume of enclosed air	Pressure exerted	Product of volume and pressure
V	A.P.	$V \times \text{A.P.}$
V_1	$\text{A.P.} + h_1$	$V_1 \times (\text{A.P.} + h_1)$
V_2	$\text{A.P.} + h_2$	$V_2 \times (\text{A.P.} + h_2)$
V_3	$\text{A.P.} + h_3$	$V_3 \times (\text{A.P.} + h_3)$
etc.	etc.	etc.

What can be said (within the limits of experimental error) about the products of the volume and pressure?

What is the percentage of error (variation) in the experiment?

5. If the above experiment were carried out using *any gas*, other than air, corresponding results would be obtained. Knowing this, and assuming that no temperature change occurs during the experiment:

Make a generalization that involves volume, pressure, and temperature for a given mass of any gas. What is this law called?

EXPERIMENT 46. (Demonstration) To determine the coefficient of expansion (volumetric) of air at a definite temperature. (Charles's Law)

1. The apparatus (see Fig. 44) consists of a clean dry 100 ml. flask, a 200 ml. beaker, thermometer, retort stand, gauze, burner, one-hole stopper, short pieces of glass and rubber tubing, clamp, and a sheet-metal cover with a semi-circular hole of suitable size to accommodate the neck of the flask.

Insert the short glass tube in the stopper and attach the piece of rubber tubing, making sure that there are no leaks. Insert the rubber stopper in the flask and place the flask in the beaker which contains water to a depth of about one inch. Place the clamp loosely over the rubber tubing. Cover the beaker with the sheet metal and heat until the water has been boiling for about five minutes (Fig. 44A).

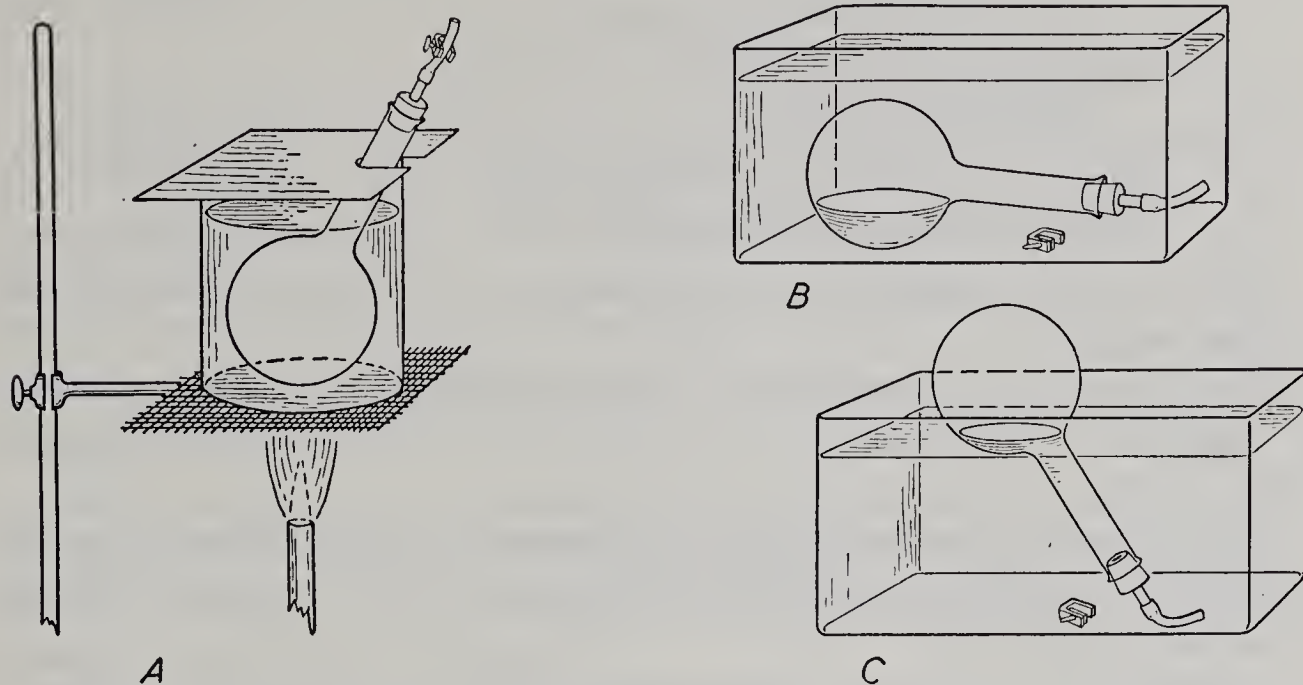


Fig. 44. Determining the coefficient of expansion (volumetric) of air at a particular temperature.

2. Tighten the clamp, thus enclosing a volume of heated air. Invert the flask in cold water in the sink or a pneumatic trough. Make sure that the stopper and tube are kept covered with water. Remove the clamp and hold the flask under water to cool the air, and until no more water enters it (Fig. 44B).

Equalize the levels of the water inside and outside the flask to ensure that the air pressure inside is balanced by the atmospheric pressure outside (Fig. 44C). Clamp the tube under water at the same place it was formerly clamped. Remove the flask from the sink and turn it right side up on the table. Remove the clamp and loosen the stopper in order to allow any water that may be trapped in the tubing to run into the flask. Remove the stopper and take the temperature inside the flask close to the contained water. Record this temperature in degrees C.

Pour the water from the flask into a graduated cylinder and record the volume of the water that entered the flask.

Fill the flask with water and insert the stopper to the same depth as before. Clamp the rubber tubing at the same position as it was

clamped previously and flip any water from the tubing above the clamp. Remove the clamp, and then the stopper, allowing any water contained in the glass and rubber tubings to fall back into the flask. Measure the volume of water in the flask.

Record the observations as follows:

1. Temperature of the boiling water
(approximately) 100°C .
2. Temperature of the air after cooling
(resulting temp.) $^{\circ}\text{C}$.
3. Volume of water that entered flask ml.
4. Volume of water needed to fill flask ml.

Why is 100°C . merely an approximation of the temperature of the boiling water?

Explain why the volume of water needed to fill the flask (Observation 4) represents the volume of air at 100°C . that was cooled during the experiment.

Explain why the volume of water that entered the flask (Observation 3) represents the shrinkage in volume of the air at 100°C . when cooled to the resulting temperature.

Calculate the coefficient of expansion of air at the temperature (Observation 2) recorded above. (For assistance in these calculations consult pages 93 and 94 of the text.)

QUESTIONS

1. Given that the coefficient of expansion of air at 0°C . is $\frac{1}{273}$, explain why the coefficient of expansion of air at 10°C . is $\frac{1}{283}$. (TEXT p. 94)

2. Explain how Lord Kelvin developed the *absolute* temperature scale.

3. You are told that all gases have the same coefficient of expansion for any given temperature, and also have noted that no pressure changes occurred during the above experiment. Make a generalization that involves volume, absolute temperature, and pressure, for a given mass of any gas. What is this law called?

UNIT 8

Chemical Laws and Terms

✓ EXPERIMENT 47. The Law of Conservation of Mass.

In this experiment a chemical reaction is performed in an enclosed space so that the reacting substances and the products of the reaction can be weighed. The class should be divided into sections, each section being assigned the investigation of one reaction only. The substances used should be examined as *solids* before their concentrated water solutions are distributed. A class discussion involving all the reactions investigated by the various sections should logically follow.

The following pairs of *solutions* are suggested: sodium chloride and silver nitrate; lead nitrate and potassium iodide; copper sulphate and sodium hydroxide; ferric chloride and potassium hydroxide; sodium sulphate and barium chloride.

Solutions of sodium chloride and silver nitrate are used in the description that follows. The same experimental procedure applies when different pairs of substances are used.

1. The apparatus used consists of a flask and a small test-tube which will rest inside the flask, but cannot fall over (Fig. 45).

Half fill the test-tube with the silver nitrate solution. Pour about 25 ml. of sodium chloride solution into the flask and then carefully place the test-tube and contents in an upright position within the flask. Insert a rubber stopper firmly in the flask and weigh the stoppered flask and its contents (Fig. 46A).

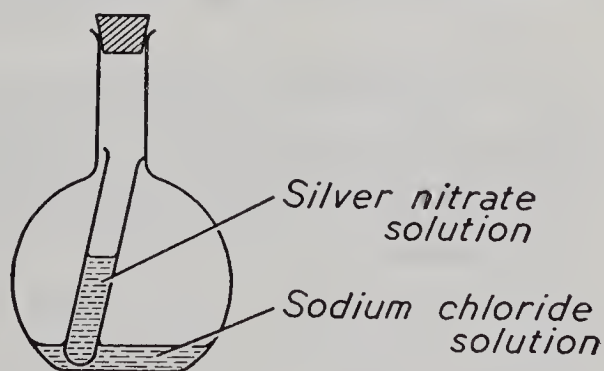


Fig. 45.

2. Leave the weights in place on the scale pan of the balance used. Remove the stoppered flask from the balance; tilt it to mix the reagents. Immediately replace the flask on the balance pan and reweigh. Remove the flask from the balance and examine the contents of the flask.

Describe the appearance of the solids used in making the solutions and the appearance of their water solutions.

What evidence is there that a chemical change has taken place?

Name the new substances that are formed. Identify the precipitate formed.

Has there been any change in weight as a result of the reaction?

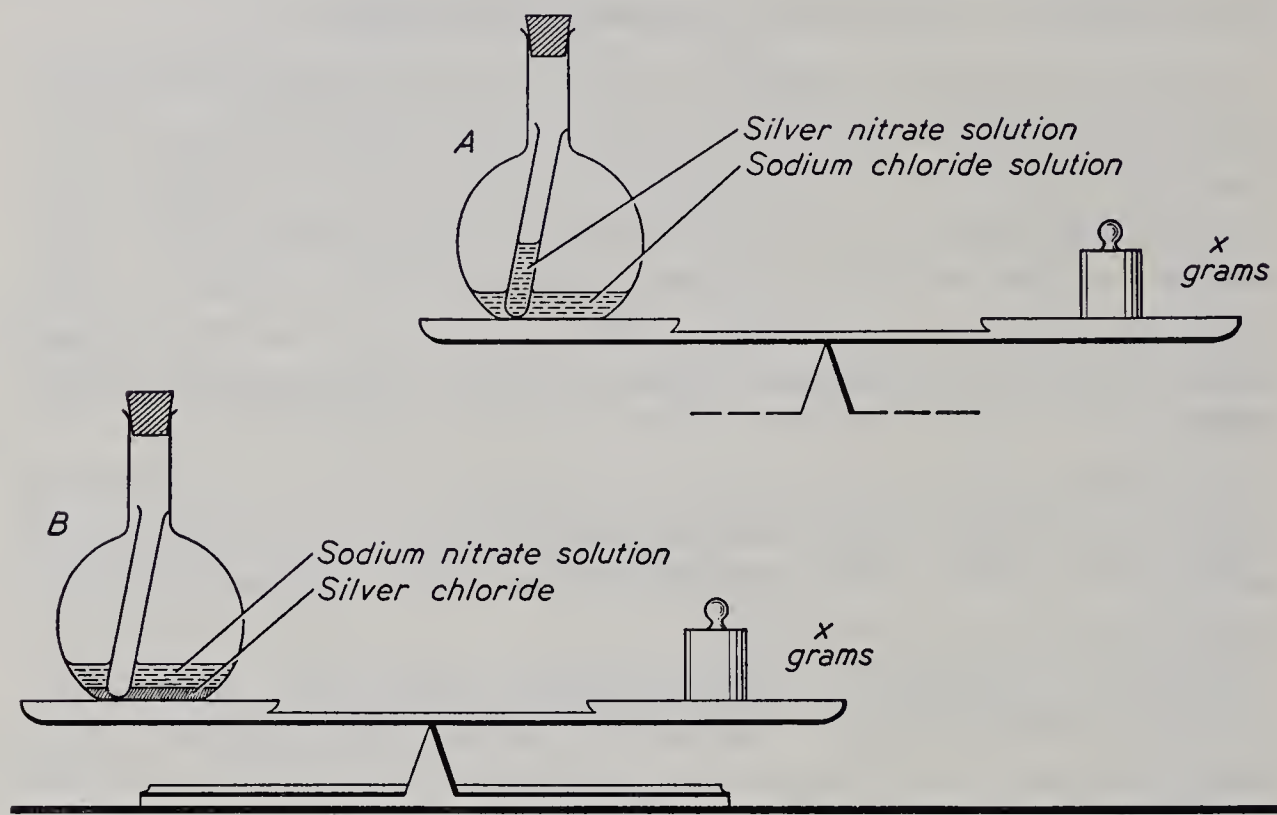


Fig. 46. The Law of Conservation of Mass.

3. Copy from the blackboard the results of the experiments carried out by other members of the class and study these results.

What two results are common to all these experiments?

4. Assume that if many thousands of chemical reactions were studied, results similar to those found in the class experiments would apply.

State a general principle or law that incorporates all chemical reactions. Name this law.

✓ EXPERIMENT 48. The Law of Definite Proportions.

1. In Experiment 8 it was found that magnesium (or copper) increased in weight when heated in air. Re-examine the observations for this experiment and calculate the percentage composition of the magnesium oxide (or copper oxide) synthesized. Compare your results with those of the class who were working with the same synthesis. (These results should be listed on the blackboard.)

Why is the reaction known as a synthesis?

What is the percentage of (a) magnesium, (b) oxygen in magnesium oxide?

How does your result compare with those obtained by other members of the class?

Within the limits of experimental error, what may be stated about the class results?

2. Dry a clean Pyrex test-tube by heating it in a Bunsen flame. Allow it to cool and weigh it accurately. Add about 1 gm. of mercuric oxide making sure that none of the powder touches the sides of the tube (TECH. 8), and weigh again. Hold the tube in a test-tube holder at an angle of about 15° to the horizontal, and heat strongly at the closed end until the mercuric oxide is all decomposed. (Keep the flame directly beneath the mercuric oxide and avoid overheating the sides of the tube.) Allow the tube to cool to room temperature and weigh the tube and its contents again.

Record the results as follows:

1. Weight of empty tube. _____gm.
2. Weight of tube and mercuric oxide. _____gm.
3. Weight of tube and mercury. _____gm.
4. Weight of mercuric oxide (2-1) _____gm.
5. Weight of mercury obtained (3-1) _____gm.
6. Weight of oxygen given off (4-5) _____gm.

Calculate the percentage composition of mercuric oxide.

What is the authority for saying that the weight of oxygen produced is equal to the weight of the mercuric oxide used minus the weight of the mercury obtained?

3. Record (from the blackboard) the results obtained by other members of the class.

How do the class results compare? Allowing for experimental error what might be stated about these results?

Make a statement about the ratio by weight of the constituents of mercuric oxide.

List all the sources of error in this experiment.

Why is it not advisable to attempt to decompose the mercuric oxide that forms again on the sides of the tube just below the film of mercury? What caused the formation of the oxide at this place?

QUESTIONS

1. If time had permitted, you would have been asked to heat the empty test-tube used in the mercuric oxide experiment "to constant weight", before adding the oxide. How would you do this? What effect would this have had on your result?

2. Why is the experiment involving magnesium oxide called a *synthesis*, and that with mercuric oxide an *analysis*?

3. In addition to magnesium oxide and mercuric oxide, many compounds have been both synthesized and analysed and in every case the proportion by weight of the constituents that make up a given compound has shown no variation. State the law that enunciates this principle. What is this law called?

✓ EXPERIMENT 49. **The distinction between a mixture and a compound.**

1. Review the properties of sulphur and iron as determined in Experiments 2 and 4.

List the properties of sulphur and those of iron.

2. Weigh on a piece of paper about 10 gm. of fine degreased iron filings, and on another paper 6 gm. of powdered sulphur. Thoroughly mix the iron and sulphur. Examine the mixture with a magnifying glass. (Iron filings can be degreased by washing them several times in alcohol until the decanted liquid is almost colourless. The filings are then washed with ether and allowed to dry by exposure to air or by being warmed gently.)

What is the colour of the mixture?

Can you distinguish the iron from the sulphur?

3. **CAUTION: Keep flames away. Carbon disulphide is volatile and very inflammable.**

Put about *one-fifth* of the mixture in a test-tube and add about 2 ml. of carbon disulphide. Shake well and filter.

Catch the filtrate on a watch-glass and allow it to evaporate. Examine the residue on the filter paper. Also examine the residue left on the watch-glass after the evaporation of the carbon disulphide.

Describe the residue on the filter paper. How could you test this residue so as to identify it positively?

Describe the residue on the watch-glass and identify it.

What property of sulphur and what property of iron is shown by this experiment?

4. Place about one-fifth of the original mixture on a piece of paper and draw a magnet through it.

Describe what happens. Can the iron and sulphur be separated by this method?

5. Drop a pinch of the mixture into a beaker of water, stir and observe what happens.

State the reason for the difference in the behaviour of the sulphur and the iron.

6. Drop about one-third of the mixture remaining into a test-tube containing about 2 ml. of dilute hydrochloric acid. Apply a lighted match to the mouth of the test-tube.

Is there any evidence of a chemical reaction?

Has the gas produced any odour? Does the gas burn? Identify it.

What substance is left in the test-tube? Why?

7. CAUTION: The reaction may be violent.

Place the remainder of the mixture in a test-tube and heat gently until the mixture glows. Then immediately withdraw the flame.

Is there any evidence that a chemical change is taking place? Describe.

8. When the substance in the test-tube has cooled, break the tube in a mortar and remove the contents. Remove the broken glass and crush the contents in the mortar with a pestle. Examine the powder with a lens.

What is the colour of the substance? Can you visibly distinguish the iron and sulphur?

9. Remove about one-fifth of the material from the mortar and shake it with 2 ml. of carbon disulphide in a test-tube. Filter, collect the filtrate on a watch-glass, and allow it to evaporate.

Is there any residue? Is the substance sulphur?

10. Remove about half of the material remaining in the mortar and draw a magnet through it.

Is the substance magnetic? Is it iron?

11. Place the remainder of the substance from the mortar into a test-tube and add about 2 ml. of dilute hydrochloric acid.

Is there any evidence of a chemical reaction?

Describe the odour of the gas produced. Is it the same gas as that formed when the mixture of iron and sulphur was dropped into hydrochloric acid? Name the gas. (TEXT p. 200)

The sulphur and the iron have combined to form a compound which is called ferrous sulphide.

QUESTIONS

1. Give four characteristics of the mixture of iron and sulphur, and four characteristics of the compound ferrous sulphide.
2. Was the reaction that took place when the iron and sulphur combined endothermic or exothermic? State the evidence for your conclusion.
3. Consult the text p. 85, to find an additional difference (not shown by this experiment) between a mixture of iron and sulphur and a chemical compound of iron and sulphur.

EXPERIMENT 50. To find the gram-equivalent weight of magnesium.

1. Cut off an 8-inch strip from a roll of magnesium ribbon. Clean the metal by rubbing it with emery paper or fine sandpaper. Roll it

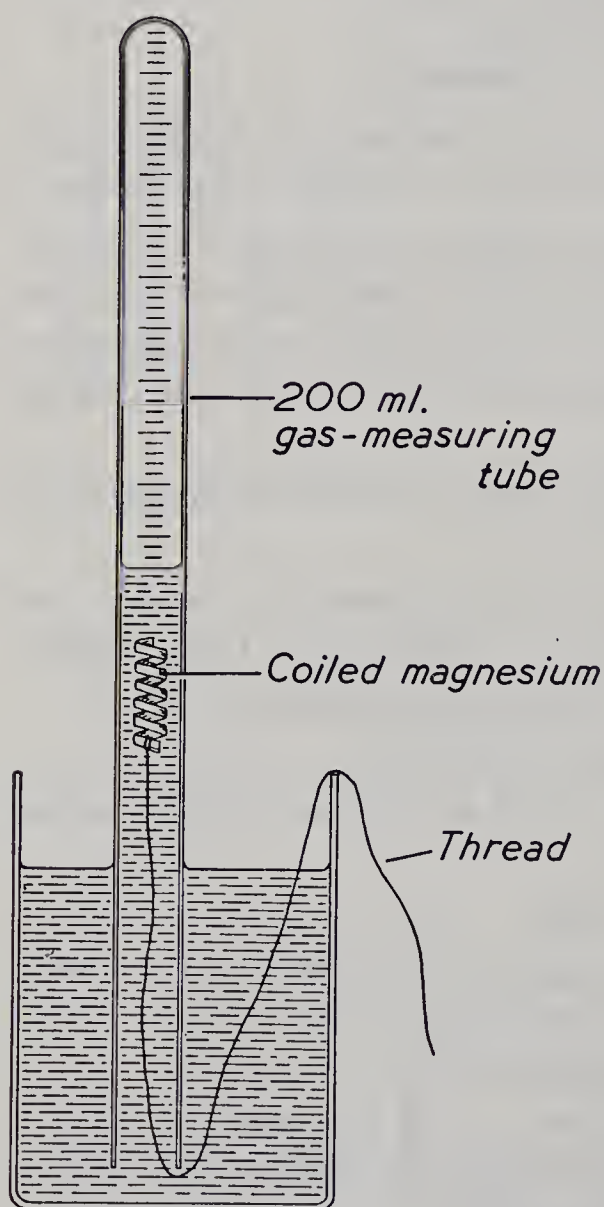


Fig. 47. Determining the gram-equivalent weight of magnesium.

into a coil small enough to slide easily into a gas-measuring tube (capacity 200 ml.) and weigh it. Attach a piece of thread to the coil leaving about two feet of thread free. (If a 100 ml. graduated cylinder is used instead of the gas-measuring tube, use a piece of magnesium ribbon 4 inches in length.)

2. Three-quarters fill a large battery jar or deep sink with water. Place about 20 ml. of concentrated hydrochloric acid in the gas-measuring tube (or 10 ml. if a 100 ml. cylinder is used) and then slowly fill it with water by pouring the water along the sloping tube to avoid, as much as possible, mixing the acid and water. Place the palm of the hand over the open end of the tube and invert it in the jar of water. Remove the hand. Immediately push the weighed coil of magnesium up into the tube and support it there with a finger until it begins to rise in the measuring tube. If the magnesium

should stick to the side of the tube a pull on the thread will again return it to the acid solution.

Why does the magnesium ribbon rise in the tube?

What gas is being collected? What is the source of this gas?

3. When all action has ceased, cool the tube by pouring water from the jar over it for five minutes. Adjust the measuring tube so that the surface of the liquid inside the tube is at the same level as the surface of the liquid in the jar. Carefully read and record the volume of the hydrogen collected.

4. Record the temperature of the liquid in the jar.

5. Record the barometric pressure of the atmosphere.

Tabulate all observations as follows:

1. Weight of magnesium used _____ gm.
2. Volume of hydrogen collected _____ ml.
3. Temperature of hydrogen _____ °C.
4. Barometric pressure _____ mm.
5. Pressure due to water vapour
(See table p. 174) _____ mm.
6. Pressure exerted by hydrogen (4–5) _____ mm.

The pressure exerted by the atmosphere (barometric reading) is equal to the pressure exerted by the hydrogen plus the pressure exerted by water vapour which is mixed with the hydrogen. For example if the barometric reading is 754 mm. and the water vapour pressure is 12 mm., then the pressure exerted by the hydrogen is $(754 - 12) = 742$ mm.

6. Calculations:

Volume of hydrogen at _____ °C. and _____ mm. _____ ml.

Volume of hydrogen at S.T.P. _____ ml.

Given that 22,400 ml. of hydrogen at S.T.P. weigh 2 gm., calculate the weight in grams of the hydrogen collected.

Calculate the weight of magnesium required to liberate 1 gram of hydrogen from the acid.

This weight is the *gram-equivalent weight* of magnesium.

7. List the class results and take the average.

What is the percentage of error in the class average?

List all possible sources of error.

EXPERIMENT 51. To find the gram-molecular weight of oxygen.

The gram-molecular weight of a gaseous pure substance is the weight in grams of that substance that will occupy 22.4 litres at S.T.P.

1. Mix about 3 gm. of dry manganese dioxide and 6 gm. of potassium chlorate on a piece of paper and put the mixture into a dry Pyrex test-tube. Insert a loose plug of glass wool above the mixture and insert a stopper containing a short piece of glass tubing in the test-tube. Weigh this generator and its contents. Arrange the apparatus as shown in Fig. 48. The pneumatic trough should be deep and the capacity of the collecting bottle should be about 2 litres.

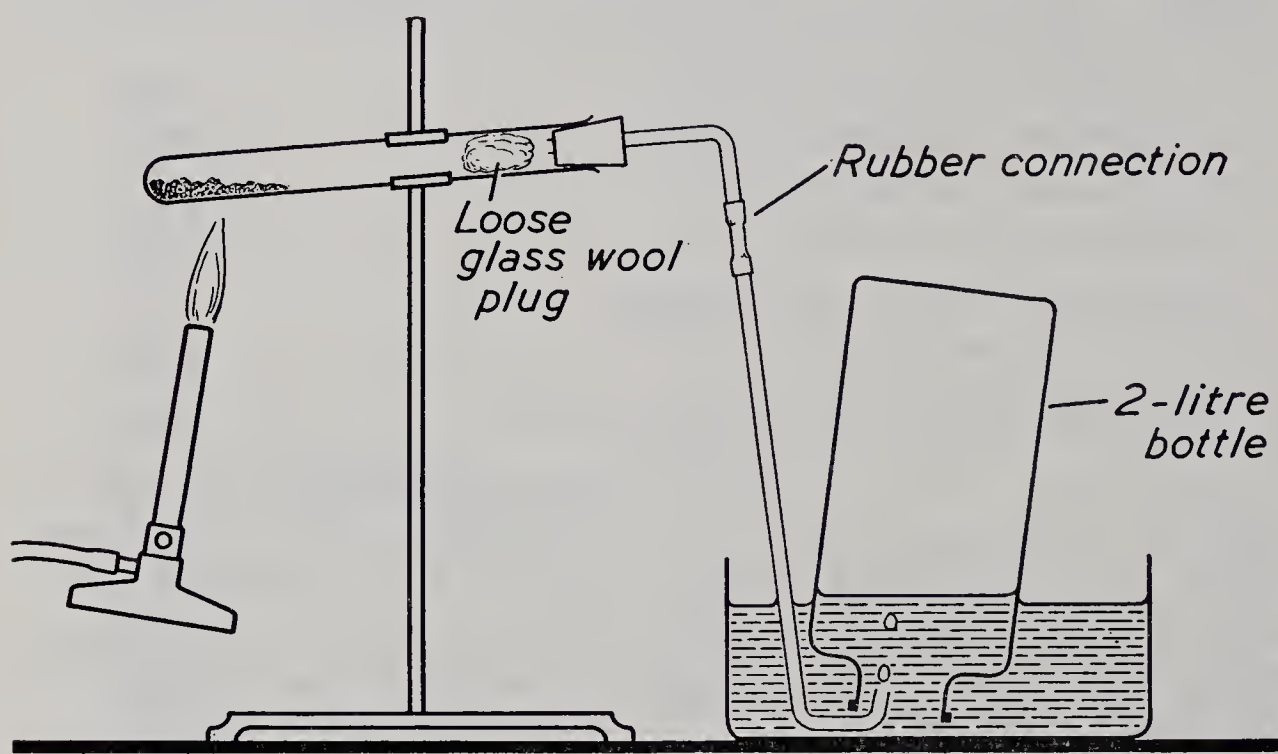


Fig. 48. Determining the gram-molecular weight of oxygen.

2. Fill the collecting bottle with water and invert it in the trough. Place the end of the delivery tube inside the bottle and heat the generator tube carefully. The heating should be controlled so that the oxygen will be evolved slowly. You should be able to count the bubbles as they rise in the collecting bottle. Continue heating the mixture until the collecting bottle is about three-quarters filled with gas. Stop heating, and when the bubbles stop rising in the collecting bottle, immediately remove the delivery tube from the pneumatic trough.

3. Adjust the level of the water inside the collecting bottle until it is the same as the level of the water in the trough. Maintaining

this level, cover the mouth of the bottle with a glass plate and then turn the bottle right side up on the desk top. Remove the cover and, using a graduated cylinder, add sufficient water to fill the collecting bottle. Keep a record of the volume of water required to fill the bottle. This represents the volume of oxygen produced. *Why?*

4. Record the temperature of the water in the pneumatic trough. Also record the barometric pressure of the atmosphere.

5. Reweigh the oxygen generator when it has cooled.

6. Record the observations as follows:

1. Weight of generator and contents before heating _____ gm.
 2. Weight of generator and contents after heating _____ gm.
 3. Weight of oxygen liberated (1–2) _____ gm.
 4. Temperature of oxygen (temp. of water in trough) _____ °C.
 5. Barometer reading _____ mm.
 6. Pressure due to water vapour (See table p. 174) _____ mm.
 7. Pressure of oxygen (5 – 6) _____ mm.
 8. Volume of gas collected _____ ml.
 9. Volume of oxygen at S.T.P. (calculated) _____ ml.
 10. Weight of 1 litre of oxygen at S.T.P. (calculated) _____ gm.
 11. Weight of 22.4 litres of oxygen at S.T.P. _____ gm.
7. The *gram-molecular weight* of oxygen is gm.

QUESTIONS

1. What was the purpose of using a loose glass wool plug in the above experiment?

2. Why was the generator heated carefully?

3. Why were the temperature and pressure of the gas collected measured?

4. List all sources of error in the above experiment.

5. State Boyle's Law and Charles's Law and explain how both laws are applied in calculating the volume of gas collected at S.T.P.

6. What is the gram-molecular weight of oxygen? What is its molecular weight? What is its molecular formula?

7. Why were the first bubbles, which obviously were air, collected in this experiment, rather than being discarded as they were in the laboratory preparation and collection of oxygen? Is this a serious source of error? Explain.

EXPERIMENT 52. To find the gram-molecular weight of carbon dioxide.

1. Record the room temperature and the barometer reading.
2. Place a rubber stopper in a clean, dry 2-litre flask, and weigh the stoppered flask. Record the weight.

Given that 1 litre of air at S.T.P. weighs 1.29 grams, calculate the weight of 2 litres of air at room temperature and at a pressure indicated by the barometer reading.

What would the stoppered flask weigh if all the air were removed?

3. Connect a piece of rubber tubing to a cylinder of carbon dioxide and insert a piece of glass tubing about a foot long into the rubber tubing. Remove the rubber stopper from the 2-litre flask and insert the glass tubing into the flask so that it almost touches the bottom. Open the valve on the cylinder of carbon dioxide and allow carbon dioxide to escape, under pressure, into the flask for about 10 seconds. By this time practically all the air within the flask will be replaced with carbon dioxide. Insert the stopper and again weigh the flask. Record this weight.

What is the weight of the stoppered flask filled with carbon dioxide?

Calculate the weight of two litres of carbon dioxide at room temperature and pressure. What would this volume be at S.T.P.?

Calculate the weight of 22.4 litres of carbon dioxide at S.T.P.

What is the gram-molecular weight of carbon dioxide as determined by this experiment? What is the correct gram-molecular weight of carbon dioxide? (TEXT p. 108) What is the percentage of error?

What is the molecular weight of carbon dioxide?

EXPERIMENT 53. To find the valence of a metal by the displacement of hydrogen from an acid.

1. The experimental method is the same as that used in Experiment 50. The average of the class results in Experiment 50 may be taken as a starting place for determining the valence of magnesium, or the experiment may be repeated. The gram-atomic weight of magnesium is taken as 24 grams.

What is meant by saying that the gram-atomic weight of magnesium is 24 grams?

What was the weight of magnesium used to displace hydrogen from dilute hydrochloric acid?

What was the calculated volume of hydrogen collected at S.T.P. displaced by this weight of magnesium?

2. Calculate (to the nearest whole number) the weight in grams of hydrogen that would have been liberated from hydrochloric acid if 24 grams of magnesium had been used. This *number* (i.e. the number of grams of hydrogen displaced) is the valence of magnesium.

What is the valence of magnesium? Define valence.

Write the equation for the chemical reaction that takes place when magnesium is added to hydrochloric acid. How many atoms of hydrogen are liberated from hydrochloric acid by one atom of magnesium?

3. Using a hydrogen generator (with delivery tube etc.), a weighed amount of zinc, a large gas bottle, and a pneumatic trough, tell how you would find the valence of zinc. You are told that the gram-atomic weight of zinc is 65 grams.

If the teacher approves your method, proceed to find the valence of zinc.

UNIT 9

Ionization, Acids, Bases, and Salts

✓ EXPERIMENT 54. The conductivity of various solutions.

1. Nearly fill a clean 300 ml. beaker with distilled water. Submerge two copper electrodes in the water so that they do not touch, and connect them in series (Fig. 49) with a lamp and an open switch. Connect to a 110-volt source, and close the switch.

Does the filament glow? Is distilled water a conductor of electricity?

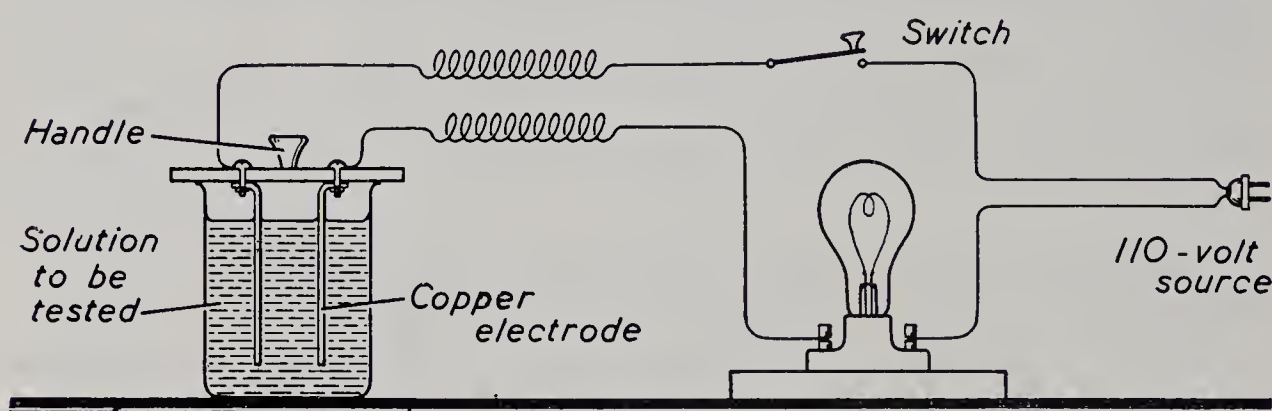


Fig. 49. Testing the conductivity of a solution.

2. Open the switch, add approximately 2 ml. of concentrated sulphuric acid to the distilled water, and stir. Close the switch and observe the lamp.

Does the filament glow? Is a water solution of sulphuric acid a conductor of electricity?

3. Pour out the dilute sulphuric acid solution and rinse the beaker and electrodes thoroughly with distilled water. Again fill the beaker with distilled water and add approximately 2 ml. of concentrated acetic acid. Stir and connect in the circuit as before. Observe the lamp.

Test the solutions made by adding approximately 2 gm. of each of the substances listed below to the beaker filled with distilled water: sodium hydroxide, ammonium hydroxide, sodium chloride, potassium nitrate, sugar, alcohol, and glycerine. (Make sure that the beaker and electrodes are thoroughly rinsed with distilled water after each test.)

List the solutions that apparently do not conduct an electric current.

Arrange the substances whose solutions conduct an electric current in the order of their conductivity, as indicated by the brightness of the lamp during the experiments. Note: For an accurate determination of conductivity, solutions of equivalent concentrations would be used.

Does the fact that the filament does not glow indicate that no electricity is passing through the circuit? Explain.

QUESTIONS

1. What three main classes of substances make good conductors of electricity when dissolved in water?
2. Use the Arrhenius' theory of ionization to explain why electrolytes conduct an electric current. (TEXT p. 153)
3. By reference to the theory of ionization explain the electrolysis of water (Expt. 35).

✓ EXPERIMENT 55. The properties of acids.

1. Put into separate, clean test-tubes half-filled with water, 5 drops of each of the following concentrated acids: hydrochloric, sulphuric, acetic, nitric, and phosphoric.

2. Put a drop of each solution on the end of a clean, glass stirring rod and taste each drop in turn.

3. Rub a drop of each solution between the thumb and the forefinger. Compare the "feel" to that of water.

4. Put a piece of both red and blue litmus paper (or one piece of neutral litmus paper) in each tube and note any changes.

Divide each of the acid solutions into four equal parts by pouring the solutions into three additional test-tubes and complete the following tests for each acid. Before continuing to the next acid in turn, rinse the three test-tubes thoroughly.

5. Add a few drops of pink phenolphthalein solution and note any change.

6. Add a few drops of neutral (green) bromthymol blue solution and note any change.

7. Add a pinch of baking soda and note any reaction. Dip a clean glass rod in limewater and hold it vertically so that a drop of limewater forms on the tip of the rod. Lower the rod into the mouth of the test-tube and observe any change in the limewater.

8. Add a pinch of magnesium powder and note any reaction. Lower a blazing splint into the test-tube and describe the result.

Complete the following table in your notebook and include *all* the observations in the seven tests carried out above. Identify the gas in all cases where a gas is produced.

NAME OF ACID USED	HYDRO- CHLORIC	SULPHURIC	ACETIC	NITRIC	PHOS- PHORIC
Formula					
Taste					
Feel					
Action on litmus					
Action on phenolphthalein					
Action on neutral bromthymol blue					
Action on baking soda					
Action on magnesium					

List all of the properties that are common to each of the acids tested.

✓ EXPERIMENT 56. The properties of bases.

1. Pour into separate clean test-tubes, to the depth of one-half inch, concentrated solutions of each of the following bases: sodium hydroxide, potassium hydroxide, ammonium hydroxide, and calcium hydroxide.

2. Add water to each test-tube until three-quarters filled and test each solution for (a) "feel" (rinse the fingers after each test), (b) taste, (c) effect on neutral litmus, (d) effect on colourless phenolphthalein, (e) effect on neutral (green) bromthymol blue, (f) effect on baking soda, (g) effect on magnesium.

3. Carry out the tests by the method suggested for acids (Expt. 55).

Construct a table similar to that used for acids and tabulate all the observations.

List all of the properties that are common to each of the bases tested.

✓ EXPERIMENT 57. Neutralization.

1. Pour 5 ml. of a dilute sodium hydroxide solution into a clean evaporating dish and add 2 drops of colourless phenolphthalein solution. Note the colour produced.

2. Add dilute hydrochloric acid to the sodium hydroxide solution with a medicine-dropper, drop by drop, stirring after each addition, until the colour just fades. Be careful not to add an excess of the acid.

Is the resulting solution basic? Is it acidic? Explain.

3. Place the evaporating dish on a wire gauze placed on a ring, clamped on a retort stand. Evaporate the solution to dryness. If the flame is removed just before the solution reaches dryness, the heat retained by the evaporating dish will dry the residue. Allow the dish to cool and examine and taste the residue.

Describe the appearance of the residue.

What is the taste of the residue? What is it?

Write an equation to represent the reaction occurring when hydrochloric acid is added to sodium hydroxide solution. Interpret this equation in terms of relative weights.

What is the name used to indicate this type of reaction? What is the significance of this term?

EXPERIMENT 58. Titrations.

When a neutralization is carried out in a quantitative way the procedure is called titration.

The instructor has prepared two solutions. One solution is sodium hydroxide and contains 40 gm. of sodium hydroxide per litre of solution. The other solution contains hydrochloric acid of a concentration somewhere between 40 gm. and 50 gm. of hydrogen chloride per litre of solution. The exact concentration is known to your instructor. Your problem is to determine the concentration of the hydrochloric acid in grams of hydrogen chloride per litre of solution.

1. Consult Technique 1 and study the structure, calibration, and the use, of a burette. Rinse two clean burettes with distilled water. (A 25 ml. pipette may be substituted for one burette.)

2. Rinse one burette with a few millilitres of the sodium/hydroxide solution. Clamp the burette on a retort stand, fill it with the sodium hydroxide solution, and adjust the level of the liquid to the zero mark (or any other recorded position). Discard any surplus solution. Rinse the second burette with the hydrochloric acid that is to be tested, clamp it on a retort stand, and fill it with the acid. Adjust the level of the liquid to the zero mark (or any other recorded position). Discard any surplus acid.

3. Place a clean beaker below the burette containing the sodium hydroxide solution and run in exactly 25 ml. of solution (or add the solution to the beaker from a 25 ml. pipette). Add 2 drops of colourless phenolphthalein solution and then place the beaker below the burette containing the hydrochloric acid.

Allow 2 ml. of the hydrochloric acid to run into the beaker and stir with a clean glass rod. Continue to add the acid in 2 ml. volumes, stirring after each addition, until the solution becomes colourless. This is a "test run" to get the approximate volume of acid solution necessary for the neutralization. Note the volume of acid required and discard the liquid in the beaker. Rinse the beaker with distilled water and repeat the titration. This time, run the acid into the sodium hydroxide solution rapidly until a volume of acid approximately 2 ml. less than that used in the test run, has been added. Then continue to add the acid drop by drop, stirring after each drop is added, until the colour just changes.

Read the burette and determine the volume of hydrochloric acid required to just neutralize the base.

Repeat the titration and find the average volume of acid used.

What was the average volume of the hydrochloric acid solution used in the titration?

Write the equation for the reaction.

What weight of hydrogen chloride is necessary to neutralize 40 gm. of sodium hydroxide? (This information is indicated by the equation for this reaction.)

If a litre of the sodium hydroxide solution contains 40 gm. of sodium hydroxide, what weight of sodium hydroxide is present in the 25 ml. of solution neutralized in the titration?

What weight of hydrogen chloride (HCl) would be needed to neutralize this weight of sodium hydroxide (i.e. the weight of the sodium hydroxide in the 25 ml. of the solution used)?

What volume of acid solution (measured in ml.) contains this weight of hydrogen chloride?

Calculate the weight of hydrogen chloride that would be present in a litre of the hydrochloric acid solution used in the titration.

UNIT 10

Carbon and its Compounds

✓ EXPERIMENT 59. Common forms of carbon.

Examine small samples of each of the following common forms of carbon: graphite, wood charcoal, soft coal, hard coal, coke, boneblack, lampblack, and carbon black. Attempt to mark a piece of paper with each form and note the "feel" of each.

Heat a small sample of graphite as intensely as possible in a Bunsen flame.

Describe the colour, hardness, feel, and general appearance of each form of carbon.

Which of these forms leaves a mark on paper?

Which form is most suitable as a writing material? Why?

Does graphite give the impression that it is crystalline in structure? Is it crystalline? (TEXT p. 166)

What is the effect of intense heat on graphite? What use of graphite does this suggest?

✓ EXPERIMENT 60. Common sources of carbon.

1. In this experiment, samples of each of the following substances are separately tested: wood, paper, cotton, wool, bread, potato, carrot, corn starch, sugar, meat, and paraffin wax.

Place a thin layer of clean quartz sand in a small crucible and on it put a small sample of the substance to be tested. Cover the crucible with a crucible cover and heat strongly for five minutes. Cool the crucible and examine the contents. For a class experiment it is suggested that each group work with a different substance.

Describe the change in appearance of the original substance in each case. What constituent is present in all the samples tested?

2. Place a small piece of each residue from Procedure 1 on a piece of mica and direct a hot Bunsen flame on each in turn.

Does the residue burn? Explain why burning did not occur in Procedure 1.

3. (a) Completely close the air inlet on a lighted Bunsen burner, and reduce the gas pressure until the flame is very small. Hold a glass plate in the tip of the flame for a short time and observe the surface of the glass.

(b) Hold a glass plate in the flame of a paraffin candle and after a short time observe the surface of the glass.

(c) Place 2 ml. of kerosene in an evaporating dish and ignite it. Hold a glass plate in the flame and after a short time observe the surface of the glass.

What is the colour of the flame in each case?

Describe the deposit formed in all three experiments. What is this deposit? What is the explanation for its formation?

QUESTIONS

1. Account for the presence of carbon in each of the substances used in the above experiment. (TEXT p. 165)

2. Explain the importance of carbon as a constituent in each of the following: wood, bread, sugar, and paraffin wax.

✓ EXPERIMENT 61. The destructive distillation of wood.

1. Arrange the apparatus as in Fig. 50 and insert either wood shavings or splints in the horizontal test-tube. Heat the test-tube with a moderate flame and observe all the changes that take place

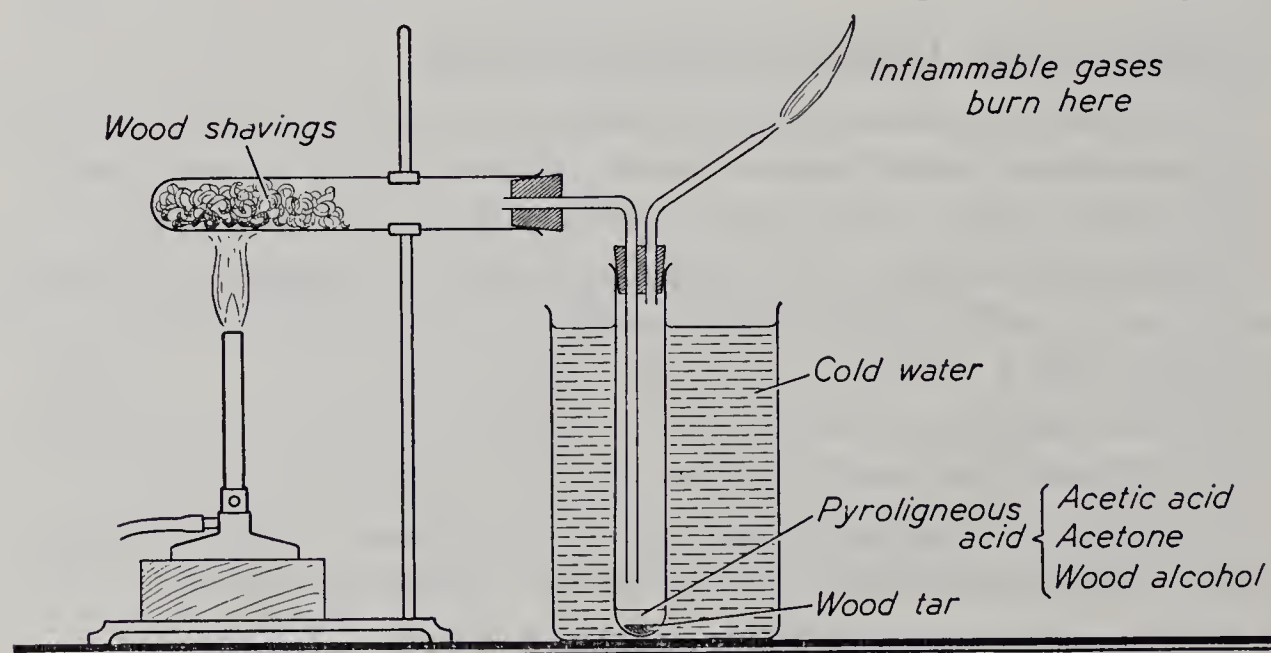


Fig. 50. The destructive distillation of wood.

within the tube. When there is evidence of a substance collecting in the vertical test-tube, bring a lighted splint to the jet and observe the colour, size, and duration of the flame.

Describe the changes that take place in the wood.

What is the nature of the deposit which forms near the mouth of the upper test-tube?

Describe the distillate which collects in the vertical test-tube.

What is the colour, size, and duration of the flame at the jet?

When does the flame go out? What is the reason for this behaviour?

What are the three distinct types of products resulting from the destructive distillation of wood?

Why did the wood not burn when heated in the test-tube?

Why is this type of distillation known as "destructive"?

2. Allow the apparatus to cool. Then remove and examine the contents of the horizontal test-tube.

Describe and name the contents. Has the wood changed throughout?

3. Note the colour, feel, and odour of the distillate. Drop small pieces of blue and red litmus papers into the liquid.

Describe the colour, feel, and odour of the distillate.

Name four substances contained in the distillate. (TEXT p. 168)

What is the effect on litmus? Which of the components of the distillate affects the litmus?

EXPERIMENT 62. The destructive distillation of coal.

1. Use the same apparatus as that used in the destructive distillation of wood (Expt. 61).

Half fill the test-tube with small pieces of soft coal. Heat the tube gently at first and then more intensely. Hold a piece of moist, neutral litmus paper at the jet and observe any change in colour. When there is evidence of a substance forming in the vertical test-tube, bring a lighted splint to the jet and observe the colour, size, and duration of the flame.

Describe and account for the change in colour of the litmus paper.

Name and give the formula of the gas responsible for this change.

Describe the changes that take place in the coal.

Describe the distillate which collects in the vertical test-tube.

What are the three distinct types of products resulting from the destructive distillation of coal?

2. Allow the apparatus to cool and remove and examine the contents of the horizontal test-tube.

Describe and name the residue. How does it compare in structure, appearance, and colour with the original soft coal?

3. Examine the distillate and note its colour, feel, and odour.

Describe the distillate as to colour, feel, and odour.

What visible proof do you have that there must be several different liquids present in the distillate?

✓ EXPERIMENT 63. **The properties of amorphous carbon.**

- 1. One-quarter fill a test-tube with activated charcoal and pour in a dilute solution of potassium permanganate to three-quarters fill the test-tube. Shake the tube for a short time and filter the contents through a moistened filter paper placed in a funnel. Collect and examine the filtrate.

What is the colour of the potassium permanganate solution?

What is the colour of the filtrate?

What important property of activated charcoal is illustrated by this experiment?

- 2. Dissolve a teaspoonful of brown sugar in 100 ml. of water in a beaker and note the colour of the solution. Add two teaspoonfuls of boneblack to the solution, and after heating for ten minutes, filter the mixture through a moistened filter paper placed in a funnel. Collect 2 ml. of the filtrate on a watch-glass. Note the colour of the solution and allow it to stand until the next laboratory lesson.

What is the colour of the original sugar solution?

What is the colour of the filtrate?

What is the nature of the residue on the watch-glass? Account for its appearance.

What important property of amorphous carbon is illustrated by this procedure?

Why is boneblack used in the refining of sugar?

3. (*Demonstration*) Pour mercury to a depth of approximately one inch into a 250 ml. beaker. Fill a test-tube with ammonia gas, insert a piece of wood charcoal into the test-tube, and invert it in the mercury. Clamp the test-tube in a vertical position and observe any change in the mercury level.

Describe and account for the change in the mercury level.

What important property of amorphous carbon is illustrated by this experiment?

QUESTIONS

1. Explain the action of activated charcoal in the canister of a gas mask. Why is activated charcoal more efficient than any other form of carbon?
2. Explain the difference between the terms *adsorption* and *absorption*.
3. Under what circumstances might a medical doctor prescribe charcoal pills?
4. What are the advantages of a charcoal dentifrice?

✓ EXPERIMENT 64. Carbon as a reducing agent.

1. Mix intimately on a piece of paper 5 gm. of cupric oxide with an equal volume of powdered charcoal, and place the mixture in a Pyrex test-tube. Fit the test-tube with a one-hole stopper and delivery tube dipping into 5 ml. of limewater in another test-tube. Heat the mixture strongly for about ten minutes and observe the reaction in each test-tube.

What evidence is there of a chemical reaction in the tube containing the cupric oxide?

Describe and account for the reaction that takes place in the limewater.

2. Allow the test-tube to cool. Empty the contents into an evaporating dish, and after examining carefully, add a few drops of concentrated nitric acid.

Describe and account for the colour of the material in the evaporating dish.

Describe and account for the action that takes place with nitric acid. (TEXT p. 257)

Write an equation expressing the reaction between cupric oxide and carbon.

Explain the action of carbon as a reducing agent.

✓ EXPERIMENT 65. Common methods of producing carbon dioxide.

1. Fill a gas bottle with oxygen (Expt. 13) and pour limewater to a depth of one-quarter inch into the bottle. Place a piece of charcoal in a deflagrating spoon, ignite the charcoal, and lower it into the bottle of oxygen. After all action has ceased, remove the spoon and shake the bottle.

Describe the behaviour of the charcoal and account for the change in the limewater.

Why is this reaction used as a test for carbon dioxide?

Write chemical equations to represent all the reactions that take place in this procedure.

2. Pour limewater to the depth of one-quarter inch into a large gas bottle and lower a burning candle into the bottle. The candle may be held with a wire or attached to the bowl of a deflagrating spoon. Remove the candle and shake the bottle.

Describe the behaviour of the candle and account for the change in the limewater.

3. One-third fill a Pyrex test-tube with magnesium carbonate and fit the test-tube with a one-hole stopper and delivery tube. Pass the delivery tube into 5 ml. of limewater in a separate test-tube. Heat the test-tube containing the magnesium carbonate and observe the limewater contained in the other tube.

Describe and account for the changes in the limewater.

Write a chemical equation for the heating of magnesium carbonate.

Repeat Procedure 3 using (a) calcium carbonate, and (b) sodium carbonate, in place of magnesium carbonate.

Does each carbonate liberate carbon dioxide when heated?

4. Into separate test-tubes place about 2 gm. of each of the following: magnesium carbonate, calcium carbonate, sodium carbonate, and sodium bicarbonate. Into the first test-tube pour 5 ml. of dilute hydrochloric acid and observe the reaction. Dip a clean glass rod in limewater and hold it vertically so that a drop of limewater forms on the tip of the rod, and bring the drop of limewater to the mouth of the test-tube. Repeat with each of the remaining three test-tubes.

What gas is given off in each case? Is the volume of gas plentiful?

Write a chemical equation to represent each reaction.

5. Repeat Procedure 4 using (a) dilute sulphuric acid, (b) dilute nitric acid, and (c) dilute acetic acid.

Is carbon dioxide given off in each case?

Write a chemical equation to represent each reaction that takes place.

6. Place one teaspoonful of baking powder in a 100 ml. beaker and add 20 ml. of water. After the reaction has subsided lower a blazing splint into the beaker.

Describe and account for the reaction that takes place when water is added to baking powder.

Describe and account for the behaviour of the burning splint.

7. Place 10 ml. of limewater in a large test-tube. Place a clean glass tube in the limewater and blow your breath through it until a change takes place in the limewater.

Describe and account for the change that takes place in the limewater.

8. (*Demonstration*) Place a teaspoonful of corn syrup or molasses, 50 ml. of water, and half a cake of yeast into a 250 ml. flask. Fit the flask with a one-hole stopper and a delivery tube that dips into 10 ml. of limewater in a test-tube. Shake the contents of the flask so as to thoroughly mix them and place the apparatus in a warm place. Observe the mixture in the flask, and the limewater in the test-tube at the end of the class period, and again the following day.

What change takes place within the flask during the class period?

Is there any change in the limewater after thirty minutes? Is there a change by the second day? Account for any change observed.

QUESTIONS

1. Name six different methods for producing carbon dioxide. Which of these methods is most suitable for producing large quantities of carbon dioxide in the laboratory?

2. Account for the presence of carbon dioxide in the atmosphere. Suggest an experiment that would prove conclusively that carbon dioxide is present in very limited amounts in the atmosphere. If the instructor approves your method, try the experiment.

✓ EXPERIMENT 66. To prepare, collect, and study the properties of carbon dioxide.

PREPARATION AND COLLECTION

1. Place two teaspoonfuls of marble chips in a 250 ml. flask, and fit the flask with a thistle-tube and a delivery tube as illustrated in Fig. 51. Add dilute hydrochloric acid through the thistle-tube until the marble chips are covered and the lower end of the thistle-tube is below the level of the acid.

2. Collect one bottle of the gas by the upward displacement of air, and three bottles by the downward displacement of water.

3. Attach a funnel to the delivery tube and allow the gas to come into contact with water in a beaker as illustrated in the diagram.

Why should the lower end of the thistle-tube be below the surface of the acid?

Which method of collection is the more satisfactory? Why?

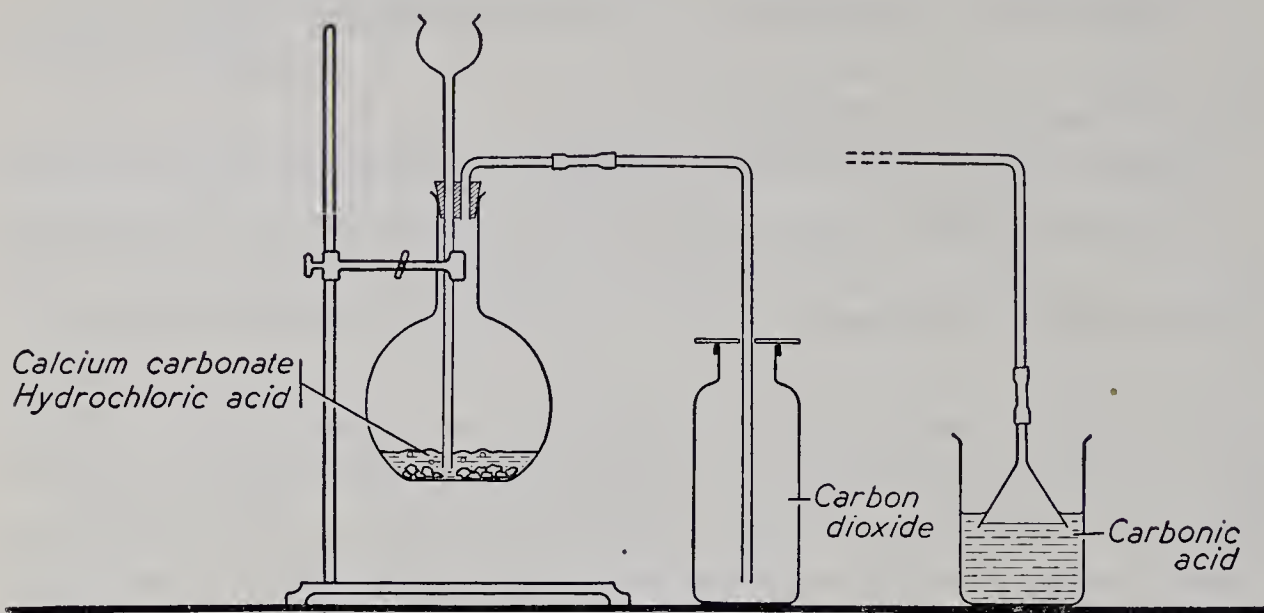


Fig. 51. The laboratory preparation and collection of carbon dioxide.

PROPERTIES OF CARBON DIOXIDE

1. Note the colour and the odour of the gas. Slowly lower a blazing splint into a bottle of the gas.

Describe the colour and odour of carbon dioxide.

Does the gas burn? Does it support the combustion of wood?

2. Wind a 6-inch strip of magnesium ribbon into a tight spiral. Hold the metal in a pair of crucible tongs and ignite the lower end of the coil. Slowly lower the burning magnesium into a bottle of carbon dioxide. Examine the deposit in the bottle.

Why does the magnesium continue to burn in the carbon dioxide?

Compare the flame with that of burning magnesium in air.

Describe the deposit formed and account for the formation of any new substances.

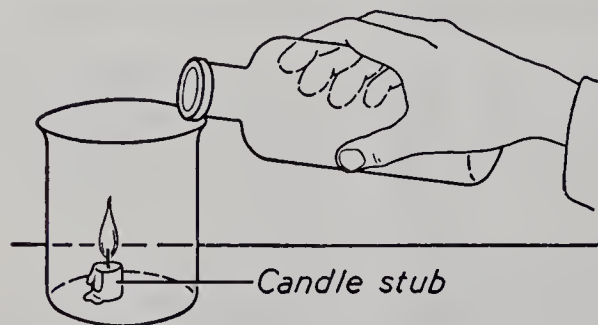


Fig. 52. Pouring carbon dioxide on a candle flame.

Write the equation to represent the reaction.

From the evidence of this experiment is it correct to say that carbon dioxide supports combustion? Explain your answer.

3. Light a small candle and stand it upright in a beaker placed on the table. Tip a bottle

of carbon dioxide over the candle flame as if pouring a liquid from the bottle (Fig. 52), and observe the action that takes place.

What happens to the candle flame? Why?

What two important properties of carbon dioxide are illustrated by this experiment? Explain.

4. Add 10 ml. of water to a bottle of carbon dioxide. Close the mouth of the bottle with the palm of your hand and shake the bottle vigorously (Fig. 53). Add another 10 ml. of water and repeat the procedure.

Add a few drops of neutral (green) bromthymol blue to the solution.

Account for the suction on your hand.

What is the effect of the solution on bromthymol blue?

What compound is present in solution?

Why was the suction on the hand less noticeable when the additional 10 ml. of water was shaken with the gas?

5. If the action in the carbon dioxide generator has ceased, add either more acid or more marble chips. One-quarter fill a test-tube with limewater and pass carbon dioxide from the generator through it until the limewater becomes milky. Continue to pass carbon dioxide through the milky mixture until it clears.

Write the chemical equation that represents the reaction that produces the milky precipitate.

Explain why the precipitate disappears and write the chemical equation that represents this reaction.

6. Drop pieces of red and blue litmus paper into the solution of carbon dioxide contained in the beaker (Fig. 51).

Describe and explain the action that takes place. Why is carbon dioxide called an acid anhydride?

7. Add 15 ml. of the carbon dioxide solution formed in the beaker to 5 ml. of limewater in a test-tube, and observe the speed with which a precipitate is formed. Note also the intensity of the precipitate.

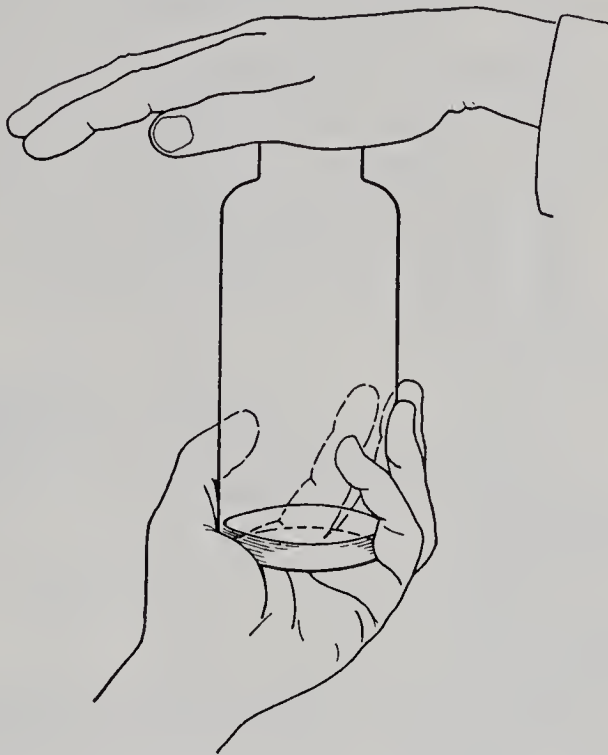


Fig. 53. Testing the solubility of a gas.

Pour 15 ml. of solution from a bottle of *soda water* into 5 ml. of limewater in a test-tube, and observe the speed with which a precipitate is formed. Note also the intensity of the precipitate.

Which has the higher concentration of carbon dioxide, the prepared solution or the "soda water"? Give reasons for your conclusions.

8. Drop a small piece of carbon dioxide "snow" into a beaker containing water and another small piece into a beaker containing lime-water. Observe each action closely.

Describe carbon dioxide "snow" or "dry ice". Why is it so called?

Describe and account for the changes observed in each of the above experiments.

QUESTIONS

1. Write the equation for the laboratory preparation of carbon dioxide.

2. List all of the properties of carbon dioxide as determined in Experiment 66.

What three properties of carbon dioxide make it useful in extinguishing fires?

✓ EXPERIMENT 67. (Demonstration) Common fire extinguishers.

1. Use a 6 oz. large-mouth bottle with vertical sides and a tight fitting screw cap. Punch a hole one-quarter inch in diameter in the

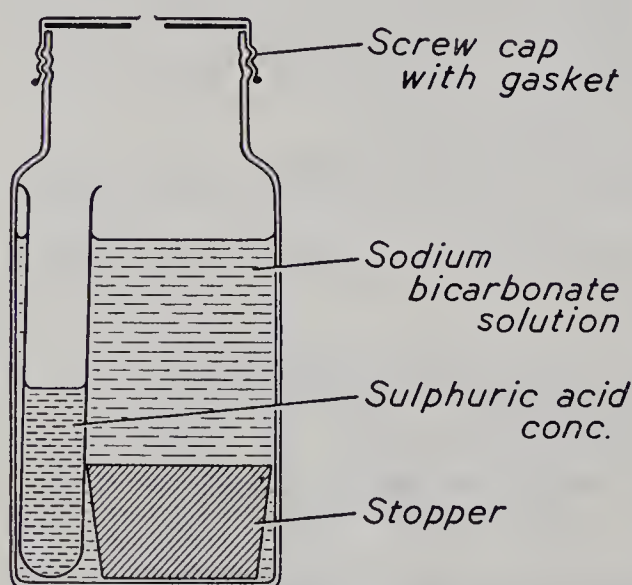


Fig. 54. A demonstration model of a soda-acid fire extinguisher.

centre of the screw cap and make sure that the paper gasket on the inside of the cap is cut away from the hole. Three-quarter fill the bottle with a concentrated solution of sodium bicarbonate. Half fill a small test-tube or vial with concentrated sulphuric acid and wedge the test-tube into the bottle with a rubber stopper so that the open end of the test-tube is one-quarter inch above the level of the baking soda solution (Fig. 54). Screw the cap on the bottle tightly and invert the bottle over the sink.

CHARACTERISTICS OF COMMON APPROVED HAND FIRE EXTINGUISHERS

Type	Plain Water*	Soda and Acid	Chemical Foam	Vaporizing Liquid		Carbon Dioxide	Dry Chemical
				Pump type	Pressure type		
Main chemicals used	Water	Sodium bicarbonate solution and sulphuric acid	Sodium bicarbonate solution and aluminum sulphate solution	Carbon tetrachloride base	Carbon tetrachloride or chlorobromomethane base	Liquefied carbon dioxide gas	Treated sodium bicarbonate powder
Expellant	Hand pump	Internal pressure from chemical reaction	Internal pressure from chemical reaction	Hand pump	Stored carbon dioxide or air pressure	Inherent pressure of liquefied gas	Carbon dioxide cartridge
Principal extinguishing agent	Water	Water	Foam	Chemical vapours	Chemical vapours	Carbon dioxide gas	Carbon dioxide gas liberated from the powder
Principal extinguishing effect	Cooling Quenching	Cooling Quenching	Blanketing Cooling	Smothering	Smothering	Smothering	Smothering
Nominal capacity	2½-5 gallons	2½ gallons	2½ gallons	1 pint 1 quart	1 quart to 2 gallons	2-20 pounds	4-30 pounds
Freezing point	32 degrees F.	32 degrees F.	32 degrees F.	50 degrees below zero F.	50 degrees below zero F.	Non-freeze	Non-freeze
Effective range	25 feet	25 feet	20-25 feet	15 feet	10-15 feet	3-8 feet	6-12 feet
Recharging material	Water	Manufacturer's charge and water	Manufacturer's charge and water	Manufacturer's extinguishing fluid only	Manufacturer's extinguishing fluid only	Manufacturer's recharge service	Manufacturer's charge and cartridge
	Class A	Yes	Yes	Surface only	Surface only	Surface only	Surface only
	Class B	No	Yes**	Small fires	Yes	Yes	Yes
	Class C	No	No	Yes	Yes	Yes	Yes

NOTE—Class A Fires—Fires in ordinary combustible materials, such as wood, paper, textile fabrics, rubbish, etc.
Class B Fires—Fires in flammable liquids, such as gasoline, kerosene, oils and greases.
Class C Fires—Fires in electrical equipment, such as switch boards, motors and wiring.

* Calcium chloride or other anti-freeze chemical may be added to plain water to depress the freezing point as required. Some are corrosive.

** Not effective on fires in alcohol, acetone, carbon disulphide, lacquer thinners, ethers or esters.
Ontario Fire Marshal's Office, Toronto, Ont.

Describe and account for the action that takes place.

Explain the action of a soda-acid fire extinguisher. (TEXT p. 176)

What class of fires may be controlled by the soda-acid fire extinguisher?

2. (*Demonstration*) Use four hydrometer jars approximately two inches in diameter and twelve inches in height. Label the jars A, B, C, and D. Pour 50 ml. of a concentrated solution of sodium bicarbonate into each jar. To jars B and D add two teaspoonfuls of licorice extract.

To A and B in turn add 50 ml. of dilute sulphuric acid and observe the action that takes place.

To C and D in turn add 50 ml. of a concentrated solution of aluminum sulphate and observe the action that takes place.

Describe the action that takes place in jar A and write the equation for the reaction.

Describe the action that takes place in jar B. What is the effect produced by the addition of licorice extract?

Describe the action that takes place in jar C and account for the colourless bubbles and the white gelatinous precipitate. Write the equation to represent the reaction.

Compare the action in jar C with that in jar A.

Describe the action that takes place in jar D. Account for the formation of the foam.

What class of fires may be controlled by a foam-type fire extinguisher?

3. Using the chart on page 115 for reference, discuss the construction, operation, and relative advantages of all types of fire extinguishers found in your school.

EXPERIMENT 68. (*Demonstration*) To prepare, collect, and study the properties of carbon monoxide.

CAUTION: Because of the poisonous nature of carbon monoxide, care must be taken to prevent the escape of more than small amounts of this gas into the room.

PREPARATION AND COLLECTION

1. Arrange the apparatus as in Fig. 55, and after the formic acid has been warmed slightly, let the concentrated sulphuric acid fall, one drop at a time, into the flask.

2. Collect one full bottle of the gas and one bottle approximately one-third full. Remove the bottles from the pneumatic trough and cover them with glass plates. Place them upright on the table until ready for use.

Describe the action that takes place when concentrated sulphuric acid is dropped into formic acid. Why is a water bath used?

Write the equation for the reaction.

What part does the sulphuric acid play in the reaction?

PROPERTIES OF CARBON MONOXIDE

1. Note the colour and odour (**CAUTION**) of the gas. Lower a lighted splint into the full bottle of the gas.

Describe the colour and odour of carbon monoxide.

Does the gas burn? Does it support combustion?

2. Pour 10 ml. of limewater into the second bottle, which contains a mixture of carbon monoxide and air, and shake. Bring a lighted splint to the mouth of the bottle. Observe the manner in which the gas burns and again shake the bottle.

Has carbon monoxide any effect on limewater?

Describe the flame produced when a mixture of carbon monoxide and air burns.

What is the product of combustion of carbon monoxide and air? Why?

Write the equation for the burning of carbon monoxide.

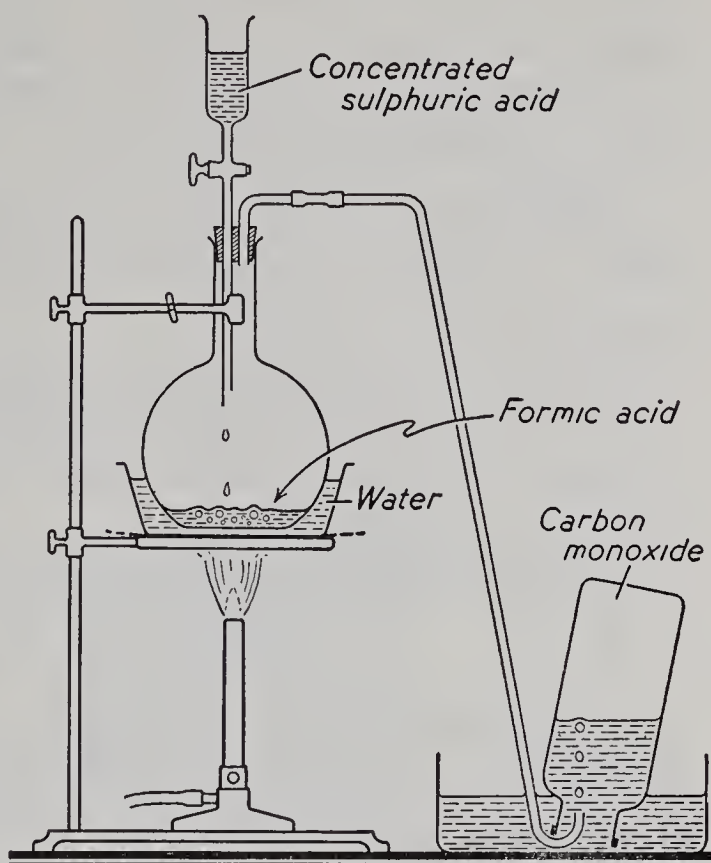


Fig. 55. The laboratory preparation and collection of carbon monoxide.

QUESTIONS

1. Why is carbon monoxide so intensely poisonous if breathed continuously? (TEXT p. 179)

2. What are the properties that cause carbon monoxide to be such an insidious gas?

✓ EXPERIMENT 69. To prepare, collect, and study the properties of acetylene.

PREPARATION AND COLLECTION

1. Half fill a 250 ml. beaker with water and invert into the beaker a test-tube filled with water. Drop a small lump of calcium carbide

into the water in the beaker. Observe all changes that take place. Invert the test-tube filled with water over the calcium carbide and collect the test-tube full of gas. Remove the test-tube and place it mouth down on the table. Collect a second test-tube half-full, another one-third full, and a fourth approximately one-twelfth full, using additional lumps of calcium carbide if necessary.

2. Lift each test-tube out of the beaker, allowing air to replace the water. Stand each test-tube mouth down on the table. Save the contents of the beaker.

Describe the appearance of calcium carbide.

Make a list of all the observations in the above experiment.

Describe the colour and the odour of the gas produced when calcium carbide is dropped into water.

Write the equation for the laboratory preparation of acetylene.

PROPERTIES OF ACETYLENE

1. Hold each of the test-tubes, one at a time, in a nearly horizontal position and bring a lighted splint to each test-tube in turn. Closely observe the flame produced. After the reaction has taken place, add 5 ml. of limewater to each test-tube and shake.

Describe and account for the way in which the acetylene burns in each case.

Why does the gas in the test-tube filled with acetylene burn so poorly?

Why is the action in the test-tube containing the least acetylene called complete combustion?

What evidence is there that complete combustion produces a higher temperature than does incomplete combustion? Explain.

What is the effect on the limewater in each case?

Write an equation to represent (a) the incomplete combustion, and (b) the complete combustion of acetylene.

2. Test the solution remaining in the beaker with a piece of neutral litmus paper. Filter some of the mixture and collect about 5 ml. of the filtrate in a test-tube. Using a clean glass tube, blow your breath through the filtrate.

What does the litmus test indicate?

What is the filtrate? Why?

What does this experiment show regarding the solubility of calcium hydroxide in water?

3. Hold the ball of the thumb firmly against the mouth of a small test-tube (preferably soft glass), so as to seal the air within, and

with a blow-pipe direct the flame from a Bunsen burner against the rounded end of the test-tube. Heat a narrow area of the tube until the expanding air within the tube forces its way through the softened glass, thus leaving a small hole in the test-tube.

Place four or five small lumps of calcium carbide in the tube. Fold or roll a filter paper into a small ball and *loosely* plug the mouth of the test-tube with the filter paper. Invert the test-tube in a beaker containing water to a depth of approximately two inches (Fig. 56). Hold a lighted match to the small opening of the test-tube.

Explain why the flame is so regular in size. Why is the term "automatic" applicable to this burner?

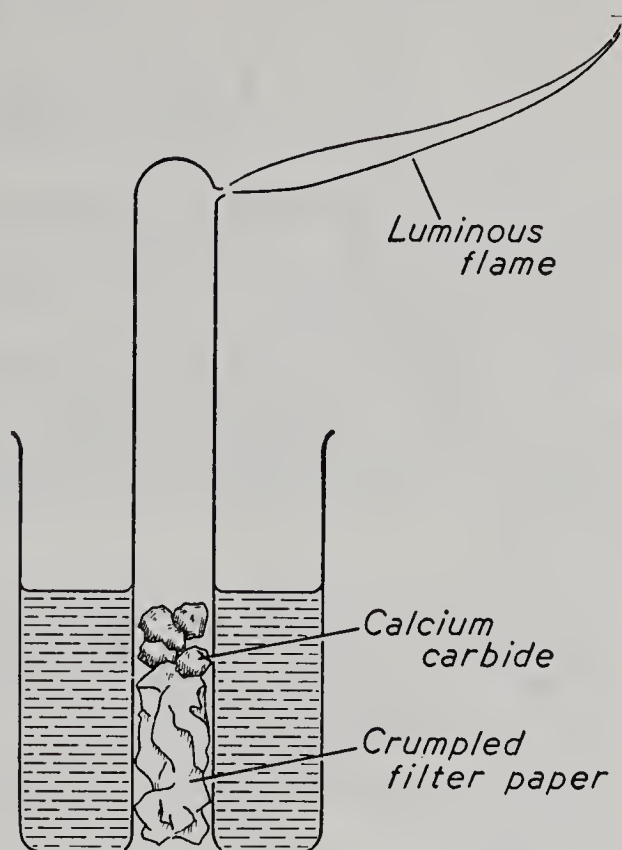


Fig. 56. An acetylene generator and burner.

QUESTIONS

1. Account for the difference in the construction of an acetylene burner used for lighting purposes, and one used for cutting and welding. (TEXT p. 181)

2. By means of suitable equations, show that decreasing the volume of oxygen in the combustion of acetylene results in an increase in the amount of carbon produced.

UNIT 11

Fuels and Flames

EXPERIMENT 70. The luminosity of flames.

1. Recall the burning of sulphur, and phosphorus in air, and in oxygen (Expt. 13).

How does the flame temperature in pure oxygen compare with the temperature of the flame in air? Explain.

How is the luminosity of a flame affected by an increase in temperature?

2. Hold (a) a piece of copper wire, and (b) a piece of glass tubing, in the flame of a Bunsen burner and observe the colour changes.

What is the effect of an increase in temperature on the luminosity of the copper and of the glass?

How is the luminosity of the flame affected by these substances? Explain.

3. Adjust the air inlet of your lighted Bunsen burner until a blue, non-luminous flame is produced. Support the burner in a horizontal position and place a large asbestos pad under the burner. Sprinkle into the flame each of the following substances, and observe the effect on the flame: (a) sodium nitrate, (b) potassium sulphate, (c) powdered calcium chloride, (d) powdered charcoal, (e) finely divided iron, and (f) zinc oxide.

Describe the effect produced on the burner flame by each substance.

In what respect is the result in (a), (b), and (c) different from the result in (d), (e), and (f)?

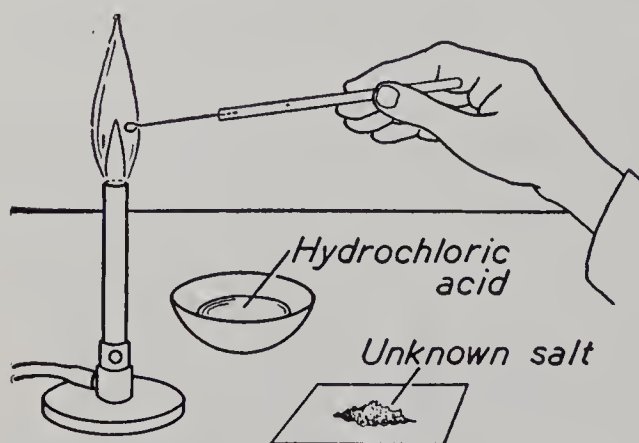


Fig. 57. Making a flame test

4. *Flame tests.* Because a Bunsen flame has a characteristic colour when some of the metals such as sodium, potassium, calcium, barium, lithium, strontium, and copper (or their salts) are heated in it, "flame tests" are sometimes used as an aid in identifying these metals. Use soluble salts of each of the metals listed above.

Adjust the Bunsen burner so that a flame as nearly colourless as possible is produced. Such a flame will be a very faint blue. Hold a platinum wire which is embedded in a glass rod (Fig. 57) in the flame. If the flame becomes coloured the wire is dirty. To clean the wire, dip it in hydrochloric acid and heat it again. Continue to dip the wire in acid and reheat until the colour of the flame is not changed. Now dip the wire in a solution of the salt to be tested, and hold it in the flame. Note any colour change of the flame. Between each test clean the wire by alternately dipping it in the acid and heating it, until no colour is imparted to the flame.

What effect has the presence of (a) sodium, (b) potassium, (c) calcium, (d) barium, (e) lithium, (f) strontium, and (g) copper on a colourless flame?

5. Close the air inlet on the lighted burner. Observe the change in the luminosity of the flame. Hold a cold glass plate horizontally in the flame for a few seconds and observe the surface of the glass.

Describe the appearance of the glass plate after it is removed from the flame. What is the deposit that forms on the plate? Account for its formation. Why does it form only when the air inlet is closed?

What substance causes the luminosity of the flame?

What happens to the solid material in the incomplete-combustion flame when the air inlet of the burner is opened?

EXPERIMENT 71. The products of the combustion of fuels.

1. Nearly fill a 250 ml. flask with cold water. Make sure that the outside of the flask is wiped dry and then hold it over the non-luminous flame of a Bunsen burner for a few seconds.

Describe the appearance of the outside of the flask.

2. Touch the deposit on the outside of the flask with a finger that has been rubbed with anhydrous copper sulphate or with a piece of dry, cobalt chloride test-paper. (This paper may be made by dipping filter paper in a cobaltous chloride solution and allowing it to dry.)

What is the liquid formed? What is the authority for your identification? Account for the formation of the liquid.

3. Over (a) a Bunsen flame, (b) a candle flame, and (c) burning kerosene in an evaporating dish, hold an inverted 250 ml. beaker which has just been rinsed with limewater.

Describe and account for the change in appearance of the film of limewater.

QUESTIONS

1. Account for the production of heat when fuels burn.
2. What precautions should be taken to obtain the maximum heat value from a fuel?
3. Under what circumstances may a fuel be used for lighting purposes?
4. Name three fuels which are also used as illuminants.
5. Name the products resulting from the *complete* combustion of fuels that contain both carbon and hydrogen.
6. Name the products resulting from the *incomplete* combustion of fuels that contain both carbon and hydrogen.
7. Write equations for the complete combustion of each of the following fuels: coke, methyl alcohol (CH_3OH), gasoline (C_8H_{10}), water gas, and acetylene.

✓ EXPERIMENT 72. The structure of flames.

1. Adjust the air supply on a Bunsen burner to produce a flame with a well-defined inner cone. Hold a piece of cardboard *horizontally* in the flame just below the tip of the inner cone. As soon as the cardboard begins to char, remove it from the flame and examine it closely.

Hold another piece of cardboard *vertically* in the centre of the flame resting the cardboard on the top of the barrel of the burner. When the cardboard begins to char, remove it from the flame and examine it.

Make separate diagrams showing the areas charred by the flame in each case.

What do the charred areas indicate regarding the temperatures of the flame? Why is the flame hotter in some areas than in others?

2. Hold the end of a piece of glass tubing, approximately 6" long, inside the smaller inner cone of the Bunsen flame, slope the tube, and bring a lighted match to the open end of the tube remote from the burner flame.

What does the result of this experiment indicate regarding the nature of the inner cone? Is the temperature inside this inner cone high or low? Why?

3. Insert a pin through a "strike-anywhere" match just below the head and place the match vertically in the barrel of an unlighted

Bunsen burner with the pin resting horizontally on the top of the barrel (Fig. 58B). Turn on the gas and light the burner.

Account for the behaviour of the match.

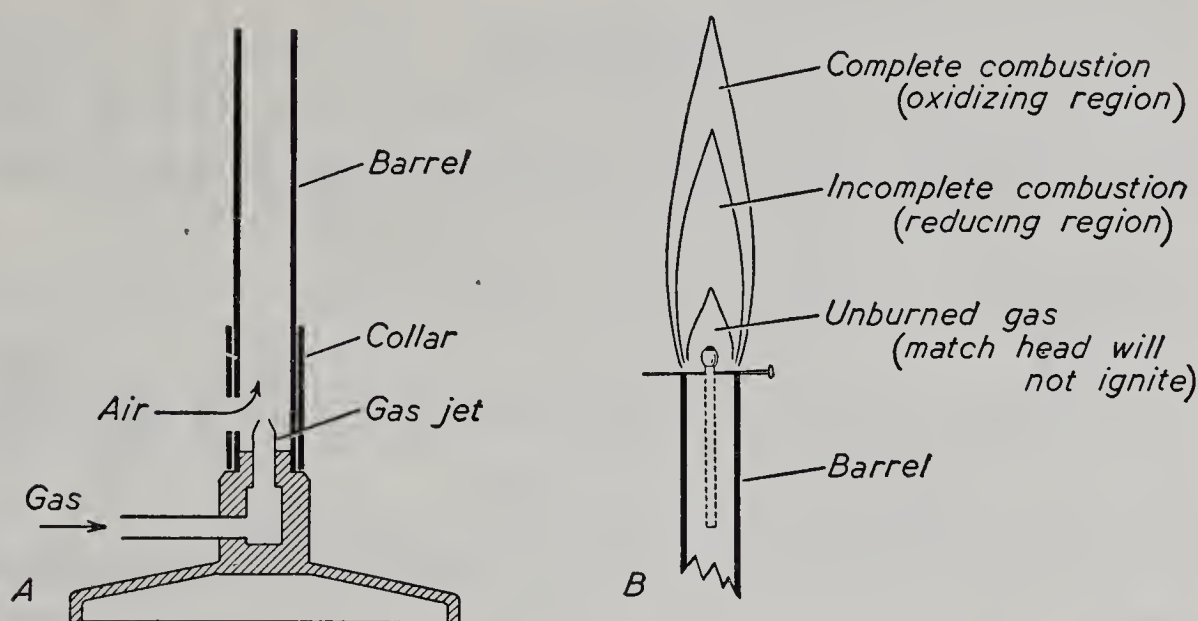


Fig. 58. A, Bunsen burner in section. B, diagram showing the various regions in a Bunsen flame.

4. Clean one end of a piece of copper wire 8" long, by dipping it in concentrated nitric acid and then washing it in water. Hold the clean end of the wire horizontally in the top portion of a Bunsen burner flame and observe the effect on the copper. Lower the wire until it is in the central cone (reducing region) and observe the colour change in the copper. Remove the wire from the flame, let the hot copper cool in air and again observe the copper.

Describe and explain the colour change that takes place when the copper is heated in the tip of the flame.

Describe and explain the colour change that takes place when the copper is lowered into the central cone.

Describe and explain the action that takes place when the hot copper is removed from the flame and allowed to cool in air.

5. Light a Bunsen burner and adjust the air supply to produce a well-defined inner cone. Hold a piece of fine copper-wire gauze horizontally over the flame and lower the gauze until it is approximately one inch above the top of the barrel. Observe the effect on the flame.

6. Hold a piece of fine copper-wire gauze horizontally 1" above an unlighted Bunsen burner, turn on the gas supply, and bring a lighted match to the gas *above* the gauze. Observe the effect of the gauze on the flame.

Describe and account for the action of the gauze on the flame in each of these experiments.

Account for the use of the Davy safety lamp. Under what conditions would it cease to be effective?

7. Bring a lighted match slowly down to the wick of an unlighted candle and observe closely the way in which the candle flame is produced. When the flame of the candle is well established, blow out the flame and immediately bring a lighted match to the wick.

Describe the action that takes place up to the time that the flame is well established.

What evidence is there that paraffin wax has a low melting point?

How does this property of paraffin make the use of a candle possible?

Why is the candle flame yellow in colour?

Why does the flame taper uniformly to a sharp point?

In what respects does a candle flame resemble the flame of a Bunsen burner?

Account for the action that takes place when the candle flame is relighted.

8. Make a helix of copper wire which will be slightly larger in diameter and slightly longer than a lighted candle flame, and bring it over the flame.

Describe and account for the action of the flame.

9. Bring a helix of copper wire, which has been heated to redness in a Bunsen flame, over a lighted candle flame.

Describe and account for the action of the flame.

QUESTIONS

1. What factors are necessary for the production of a flame?
2. Why is the tip of a Bunsen burner flame an oxidizing flame?
3. Why is the middle cone (region of incomplete combustion) a reducing region?
4. What is the function of the barrel of a Bunsen burner?

UNIT 12

Sulphur

✓ EXPERIMENT 73. The allotropic modifications and the properties of sulphur.

1. Examine a small sample of (a) roll sulphur, and (b) flowers of sulphur.

Describe the colour and appearance of (a) roll sulphur, and (b) flowers of sulphur.

2. Half fill a Pyrex test-tube with small pieces of roll sulphur. Hold the test-tube at an angle of 45° and heat the tube with a low Bunsen flame. Rotate the tube constantly and continue heating carefully until all the sulphur has just melted. Note the colour and nature of the liquid.

3. Pour the melted sulphur into a dry filter paper which has been placed in a funnel, and let it stand without jarring until there is evidence of crystals forming nearly to the centre of the sulphur surface.

Quickly pour the melted sulphur remaining in the filter paper into a beaker half full of cold water. Immediately remove the filter paper from the funnel and open the paper to expose the crystals that have formed. Examine these crystals and keep them for Procedures 4 and 5.

Describe the changes that take place in the sulphur while it is being melted.

Which is more dense, liquid sulphur or solid sulphur? Give a reason for your answer.

Does the melted sulphur wet glass?

Compare the viscosity of melted sulphur with that of water.

What is the nature of the crystals which form on the filter paper?

Name this allotropic modification of sulphur.

Describe the appearance of the sulphur formed when melted sulphur is poured into cold water.

Name this allotropic modification of sulphur.

4. **CAUTION:** Keep flames away. Carbon disulphide is volatile and very inflammable.

Place a few of the crystals that form on the filter paper (Procedure 3) in a test-tube. Add 2 ml. of carbon disulphide and shake.

Place a piece of roll sulphur about the size of a pea in another test-tube, add 2 ml. of carbon disulphide and shake.

Describe the action of carbon disulphide in each case.

Describe how you would obtain large crystals of rhombic sulphur from a solution of roll sulphur in carbon disulphide.

5. Allow the remaining crystals of sulphur on the filter paper (Procedure 3) to stand until the next laboratory period and observe any changes.

Describe and account for the changes observed.

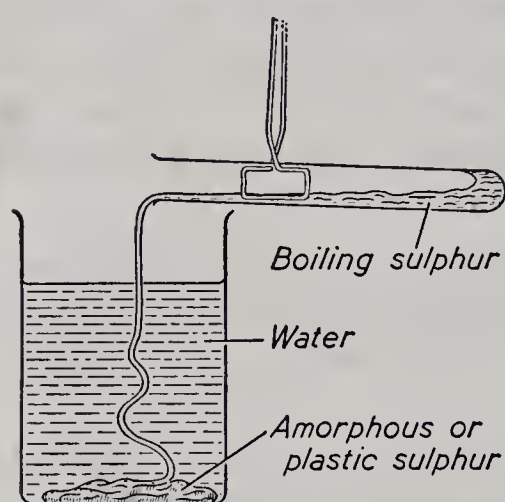


Fig. 59. Pouring boiling sulphur into cold water.

cold water (Fig. 59). Examine the material in the beaker.

Describe the changes in colour and viscosity of the sulphur from the time that the solid first melts until the liquid sulphur boils.

Give two observations that would indicate that the sulphur is boiling.

Account for the appearance, colour, and formation of the sublimate. What is its common name?

Describe the colour and nature of the sulphur resulting from pouring boiling sulphur into cold water. Name this allotropic modification of sulphur.

7. Divide the plastic sulphur, obtained by pouring boiling sulphur into cold water, into two portions. To one portion, in a test-tube, add 5 ml. of carbon disulphide and shake. Filter, collect some of the filtrate on a watch-glass and allow it to evaporate. Let the other portion of plastic sulphur stand until the next laboratory period and examine it.

6. Use the same test-tube as that used in Procedure 2 and half fill it with roll sulphur. Heat the test-tube with a low Bunsen flame until the sulphur has melted. Continue heating carefully and slowly, tilting the tube occasionally to determine the viscosity of the sulphur, until the sulphur boils. Note the successive changes in colour and in viscosity of the molten sulphur. Observe the inside of the test-tube immediately above the boiling sulphur. Slowly pour the boiling sulphur into a beaker full of

Is plastic sulphur soluble in (a) water, (b) carbon disulphide?

Describe and account for the change in appearance of the plastic sulphur when allowed to stand.

8. Place a small lump of roll sulphur about the size of a pea in an asbestos-lined deflagrating spoon and ignite the sulphur. Observe the appearance of the flame and any changes in the sulphur itself. Cautiously note the odour of the gas. Lower the burning sulphur into a bottle of air containing 10 ml. of a dilute solution of potassium permanganate. Shake and observe.

Describe the appearance of the burning sulphur.

Describe the odour of the gas produced. Name this gas.

Describe and account for the change that takes place in the potassium permanganate solution.

Write chemical equations for (a) the burning of sulphur, and (b) the action of sulphur dioxide on potassium permanganate solution. (TEXT p. 203)

Why can the action with potassium permanganate be considered a chemical test for sulphur dioxide?

9. Quarter fill the same test-tube that was used to prepare plastic sulphur, with roll sulphur. Heat the sulphur until it boils. Keep the sulphur boiling and lower either a spiral of fine copper wire or a strip of copper foil into the vapour of the boiling sulphur. Observe the changes that take place.

Describe the nature and appearance of the copper before being lowered into the test-tube.

Describe and account for the reaction that takes place in the test-tube.

What is the nature and appearance of the product resulting from the reaction?

Compare the reaction between copper and sulphur with that between iron and sulphur (Expt. 49).

QUESTIONS

1. What property of sulphur makes possible its commercial recovery by the Frasch method? (TEXT p. 196)

2. Explain the statement "sulphur exists in three different allotropic modifications".

3. How could you prove that the allotropic modifications of sulphur have almost identical chemical properties?

✓ EXPERIMENT 74. To prepare, collect, and study the properties of hydrogen sulphide.

PREPARATION AND COLLECTION

CAUTION: Hydrogen sulphide is poisonous. When inhaled in concentrations of one volume of the gas in two hundred volumes of air, it is fatal. Bleaching powder sprinkled with acetic acid provides sufficient chlorine gas to serve as an antidote.

1. Arrange the apparatus as shown in Fig. 60. Put several pieces of ferrous sulphide into the flask and add dilute hydrochloric acid through the thistle-tube until the ferrous sulphide is well covered.

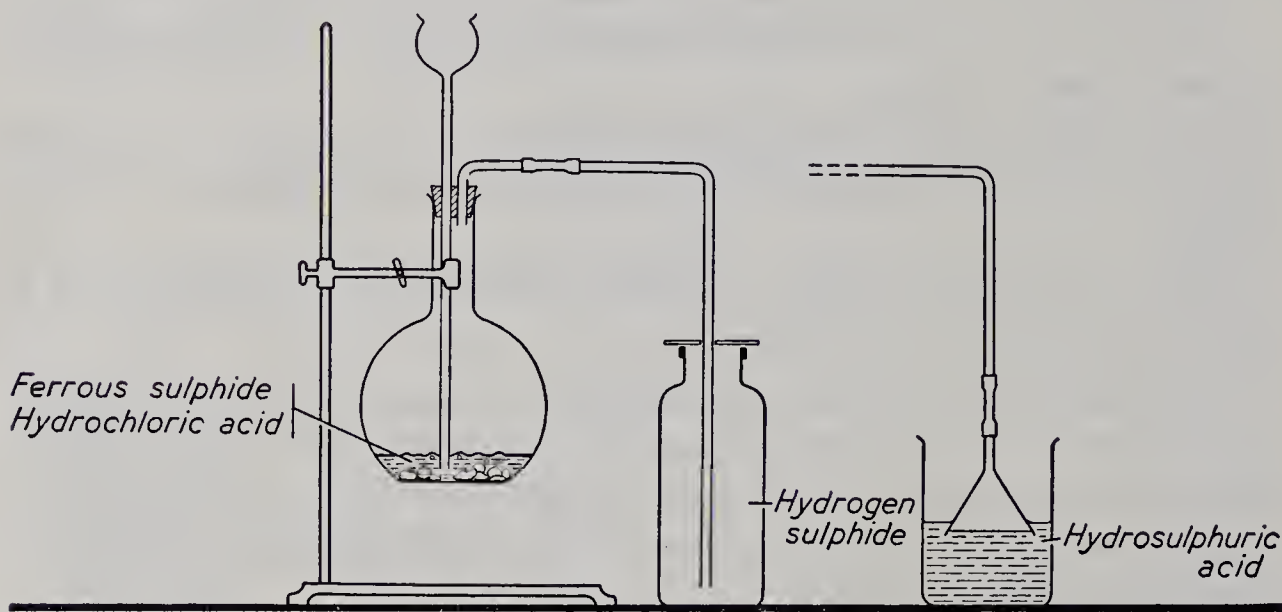


Fig. 60. The laboratory preparation and collection of hydrogen sulphide.

2. Collect five bottles full of the gas by the upward displacement of air. To determine when the bottles are filled with the gas, bring to the mouth of each bottle a piece of filter paper which has been dipped into a lead acetate solution.

3. When the gas has been collected, prepare a solution of the gas in water as indicated in the diagram. Keep this solution for future use.

Describe the action that takes place within the flask.

State two reasons for collecting the gas by the upward displacement of air.

Describe the action that takes place when hydrogen sulphide reacts with lead acetate solution. Write the chemical equation to express this reaction. Explain why this reaction serves as a test for hydrogen sulphide.

Write the equation for the laboratory preparation of hydrogen sulphide.

PROPERTIES OF HYDROGEN SULPHIDE

1. Observe the colour and the odour of the gas. Invert a bottle of the gas in a dish half full of water and allow it to stand until the end of the period.

Describe the colour and the odour of hydrogen sulphide.

What does the height of the water in the bottle indicate about the solubility of hydrogen sulphide in water?

2. Into the second bottle of gas drop (a) a piece of moist red litmus paper, (b) a piece of moist blue litmus paper, and (c) a silver coin.

Describe and account for the behaviour of each piece of litmus paper.

What change takes place in the silver coin? Write the equation for the action of hydrogen sulphide on silver.

What name is given to the solution of hydrogen sulphide in water?

3. Lower a blazing splint into the third bottle of hydrogen sulphide and observe the reaction closely.

Does hydrogen sulphide burn?

Does hydrogen sulphide support combustion?

Describe the flame of the burning gas.

Describe the substance that forms inside the bottle. What is this substance? Account for its formation and write the equation that represents this action.

4. Into a fourth bottle of the gas pour 5 ml. of a solution of lead acetate.

Describe and account for the change in appearance of the lead acetate solution. Why does the solution in the bottle have a different appearance than the deposit on a piece of filter paper which has been dipped in a lead acetate solution and exposed to the gas?

Would the solution of any other soluble lead salt react in the same way? Verify your answer by experiment.

5. Place a piece of sulphur the size of a bean in an asbestos-lined deflagrating spoon, ignite the sulphur and lower the burning sulphur into a gas bottle filled with air. Let the sulphur burn until the flame is extinguished and remove the deflagrating spoon from the bottle. Place this bottle of sulphur dioxide mouth down over the remaining bottle of hydrogen sulphide and let the bottles stand in this position for a few minutes. (If no reaction takes place, introduce a few drops of water to the mixture of gases.)

Describe and account for the action that takes place. Write the equation.

6. Test the solution of hydrogen sulphide in the beaker with a piece of neutral litmus paper. Pour some of the solution into a bottle, put a stopper in the bottle and allow it to stand until the next laboratory period.

Is the solution a strong or a weak acid? Explain.

Describe and account for the appearance of the hydrogen sulphide solution after standing for a day. Write an equation to represent the chemical change that has taken place.

7. CAUTION: Arsenic and antimony compounds are very poisonous.

Place in 6 separate test-tubes 10 ml. samples of solutions of each of the following: cupric sulphate, lead nitrate, silver nitrate, zinc sulphate, arsenic trichloride and antimony trichloride. Add in turn to each test-tube 2 ml. of freshly prepared hydrogen sulphide solution and observe the changes that take place. Observe the effect of adding 2 ml. of ammonium hydroxide to the test-tube that contained the zinc sulphate solution.

Describe and account for the changes that take place in each test-tube when hydrogen sulphide solution is added. Write a chemical equation to represent each reaction.

What is the effect of adding ammonium hydroxide to the test-tube which contains the zinc sulphide? In what way does zinc sulphide differ from the other sulphides produced in this experiment?

QUESTIONS

1. Why is the burning of a full bottle of hydrogen sulphide an example of incomplete combustion?

2. Write the chemical equation to represent (a) the complete combustion of hydrogen sulphide, and (b) the incomplete combustion of hydrogen sulphide.

3. Some scientists believe that the sulphur deposits of Louisiana have been formed by the interaction of hydrogen sulphide gas and sulphur dioxide gas. If this did take place, account for the formation of the hydrogen sulphide and the sulphur dioxide required for these vast deposits. Write a chemical equation that would indicate the reaction that took place in the formation of the sulphur.

✓ EXPERIMENT 75. To prepare, collect, and study the properties of sulphur dioxide.

PREPARATION AND COLLECTION

1. Arrange the apparatus as shown in Fig. 61 and place a concentrated solution of sodium bisulphite in the flask. Control the rate at which sulphuric acid drops into the flask so that a steady flow of gas is produced.

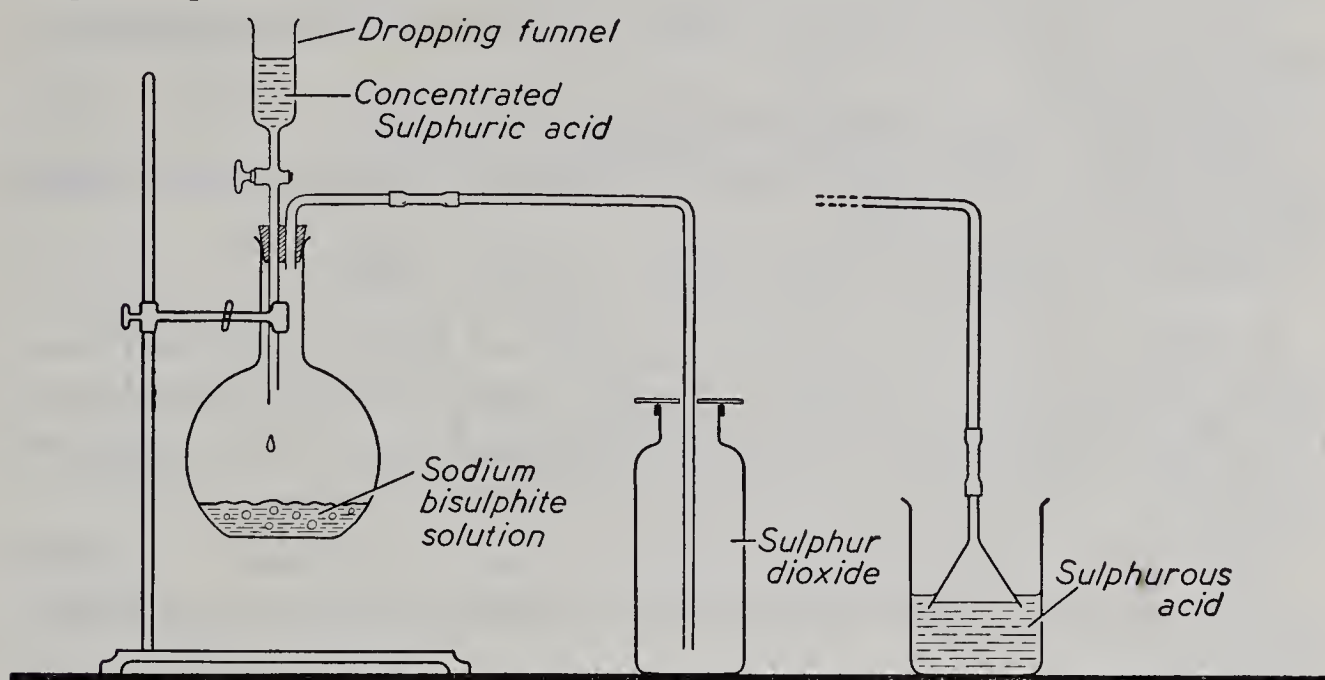


Fig. 61. The laboratory preparation and collection of sulphur dioxide.

2. Collect four bottles of the gas by the upward displacement of air and place a glass plate over the mouth of each bottle.

3. Place a funnel on the end of the delivery tube as illustrated in the diagram, and make a solution of sulphur dioxide in water.

Describe the action that takes place when the acid and the sodium bisulphite solution react.

Write the chemical equation for the laboratory preparation of sulphur dioxide.

PROPERTIES OF SULPHUR DIOXIDE

1. Observe the colour and odour of the gas. Note the effect on moist neutral litmus paper.

Describe the colour and odour of sulphur dioxide.

What change takes place in the neutral litmus paper? Account for the action of sulphur dioxide on litmus.

2. Lower a blazing wooden splint into a bottle of sulphur dioxide.

Does the gas burn? Does the gas support combustion? Explain how advantage may be taken of these properties to extinguish a fire in a chimney.

3. Pour 5 ml. of a dilute solution of potassium permanganate into the third bottle of sulphur dioxide. Add 2 ml. of barium chloride solution to the bottle and observe the effect produced. Finally, add 2 ml. of dilute hydrochloric acid to the bottle and observe the effect.

Describe and account for the change in appearance of the potassium permanganate solution. Why does this reaction serve as a chemical test for sulphur dioxide?

Write the equation for the action of potassium permanganate on sulphurous acid.

Describe the action of the barium chloride solution.

Write the equation for the action of barium chloride on the solution remaining in the gas bottle.

Why was hydrochloric acid added? (TEXT p. 208)

4. Place a thoroughly moistened, coloured flower (e.g. a red rose) in the fourth bottle of sulphur dioxide. After a short time observe the colour of the flower and then hold the flower in the vapour of concentrated nitric acid. Observe any further change.

What is the effect of sulphur dioxide on the moist flower? Explain.

What happened when the flower was held in nitric acid fumes? Explain.

Substances bleached with sulphur dioxide often return to their original colour. Explain with the aid of a suitable equation.

5. Pour 5 ml. of the sulphurous acid solution contained in the beaker into a test-tube and add a small piece of magnesium ribbon.

Describe and account for the action that takes place.

How would you identify the gas that is given off?

What product remains in the solution?

Write the chemical equation for the reaction.

6. Pour 10 ml. of the sulphurous acid solution into a test-tube, add 5 ml. of a solution of barium chloride and shake. To the contents of the test-tube add dilute hydrochloric acid, a few drops at a time, until no further change takes place.

Describe and account for the action of the barium chloride solution with sulphurous acid. Write the equation for the reaction.

Describe and account for the action when hydrochloric acid is added to the contents of the test-tube. Write the equation for the reaction.

7. Allow the remainder of the solution to stand in a beaker, exposed to the air, until the next laboratory class and repeat the test with barium chloride solution and hydrochloric acid as in Procedure 6.

Describe and account for the action of the barium chloride solution. Explain with the aid of equations.

Account for the action with hydrochloric acid.

How could you distinguish a sulphite from a sulphate?

EXPERIMENT 76. (Demonstration) A laboratory experiment to prepare sulphuric acid.

1. Select a three-hole rubber stopper to fit an 8 oz. gas-collecting bottle. Insert a short piece of glass tubing, which will act as an outlet tube, into one of the openings. Insert a glass delivery tube leading from a sulphur dioxide generator (Expt. 75) into a second opening. Insert a glass delivery tube, leading from a Pyrex test-tube containing four grams of lead nitrate, into the third opening.

2. Heat the lead nitrate until the gas bottle is filled with brown fumes; then allow the sulphuric acid to drop slowly into the flask containing the sodium bisulphite solution. Control the heating of the lead nitrate so as to always have some brown gas in the gas bottle.

3. After a short time, observe and describe the appearance of the inside of the gas bottle.

4. When brown fumes are no longer produced in the test-tube, disconnect the apparatus and slowly, a few drops at a time, add 10 ml. of water to the gas bottle.

Describe the action that takes place when water is added to the gas bottle.

5. Divide the solution produced in Procedure 4 into two parts by pouring it into two test-tubes. Add a solution of barium chloride, followed by 2 ml. of hydrochloric acid, to one test-tube. Add a small piece of mossy zinc to the other test-tube and test with a blazing splint.

Describe and account for the action with the barium chloride solution and the hydrochloric acid.

Describe and account for the action with the zinc.

QUESTIONS

1. Name the brown gas produced when lead nitrate is heated. Write the equation for the reaction.

2. What is the function of the brown fumes in the reaction? Would the oxygen produced by the heating of lead nitrate help or hinder the reaction? Explain.

3. Write the equation for the production of sulphur trioxide.
4. Describe sulphur trioxide. What happens when water is added to the sulphur trioxide in the gas bottle? Write the equation.

✓ EXPERIMENT 77. The properties of sulphuric acid.

CAUTION: Concentrated sulphuric acid is very destructive to flesh and clothing and should be handled with care.

1. (*Demonstration*) Put 40 ml. of water into a 250 ml. beaker. Place in the water a thermometer calibrated to read up to 200°C. and record the temperature. With constant stirring, pour into the water, in a fine stream, 100 ml. of concentrated sulphuric acid. Record the temperature.

Describe the action that takes place when concentrated sulphuric acid is added to water.

What temperature change is indicated by the thermometer?

Explain why water must never be added to concentrated sulphuric acid.

Why is concentrated sulphuric acid a good drying agent?

2. (*Demonstration*) Half fill a 125-ml. beaker with cane sugar ($C_{12}H_{22}O_{11}$). Add sufficient concentrated sulphuric acid to just cover the sugar and observe.

Describe the changes that take place and account for these changes.

Name the products resulting from the reaction.

Write a chemical equation to represent the reaction. What role does sulphuric acid play in the reaction?

3. Dip a wooden splint into concentrated sulphuric acid and observe the effect. Wood contains cellulose ($C_6H_{10}O_5$)_n.

Describe and account for the change that takes place.

Write a chemical equation to represent the reaction that takes place.

4. Dip a glass tube into concentrated sulphuric acid and trace a few lines with the concentrated acid on a piece of filter paper.

Describe and account for the action of the concentrated acid on the paper.

Write an equation to represent the reaction.

5. Dip a glass tube into dilute sulphuric acid and trace a few lines with the dilute acid on a piece of filter paper. Warm the filter paper over a Bunsen flame and observe the action that takes place.

What is the action of the dilute sulphuric acid on the filter paper? Why is the filter paper warmed?

6. (CAUTION) Place two drops of concentrated sulphuric acid on a piece of mica and heat the mica with a Bunsen flame.

Describe and account for the changes that take place.

7. (Demonstration) Into each of two separate test-tubes, place a piece of mossy zinc. Add 5 ml. of dilute sulphuric acid to one test-tube, and 5 ml. of concentrated sulphuric acid to the other. Warm the test-tube containing the concentrated acid.

Describe and account for the gaseous product formed in each case.

Write an equation for each reaction.

UNIT 13

Salt

EXPERIMENT 78. The solubility and crystal structure of common salt.

1. Select two test-tubes of the same size. Half fill one of them with cold water and half fill the other with hot water. To each test-tube add one level teaspoonful of common salt. Shake the tubes vigorously and allow them to stand until any undissolved salt settles.

What kind of solution is present in each test-tube?

Is there a noticeable difference in the amount of salt that remains undissolved in each case? Describe and explain.

What is the solubility of sodium chloride in water at (a) 10°C., (b) 90°C.? (TEXT p. 77)

2. Filter both solutions through a common filter paper. Collect the filtrate and divide it into two parts. By means of a Bunsen flame, rapidly evaporate one portion to dryness. Set the other portion aside to slowly evaporate to dryness. Examine the contents of each container. Select one large crystal and examine it with a magnifying glass.

What is the appearance of the residue resulting from (a) rapid evaporation, (b) slow evaporation? Account for the difference.

Describe the crystal structure of sodium chloride.

QUESTIONS

1. List four properties of common salt.
2. What property of common salt makes its commercial recovery possible?
3. Suggest a method of obtaining a large crystal of common salt.

EXPERIMENT 79. (Demonstration) The electrolysis of a solution of common salt (brine).

1. Set up the Hoffman electrolysis apparatus as shown in Fig. 39, and fill it with a saturated solution of common salt. Pass a direct

current of electricity through the solution (see Expt. 35). Closely observe the action that takes place at each electrode. When sufficient gas has collected in each arm of the apparatus, place a test-tube over each jet and collect a test-tube full of each gas.

Describe the action at the cathode.

Describe the action at the anode.

What is the colour of each gas collected?

2. Bring a lighted splint to the test-tube filled with the gas collected over the cathode.

What is the name of this gas?

3. Carefully note the odour of the gas collected over the anode and test the gas with a piece of moist blue litmus paper.

Describe the odour of the gas.

Describe the effect of the gas on moist blue litmus paper.

What is the name of this gas?

4. Place a few drops of the solution remaining in the electrolysis apparatus on a piece of red litmus paper. Test a solution of common salt with a piece of red litmus paper.

Account for the difference in the litmus tests.

What new chemical is present in the solution remaining in the electrolysis apparatus? Account for its formation.

QUESTIONS

1. Name three important substances resulting from the electrolysis of common salt solution. Explain the changes that take place to produce these substances. Write a summation equation representing the reactions that take place.

2. Name four commercial uses for each of the products formed by the electrolysis of brine.

✓ EXPERIMENT 80. To prepare, collect, and study the properties of hydrogen chloride.

PREPARATION AND COLLECTION

1. Set up the apparatus as shown in Fig. 62. Put approximately four teaspoonfuls of common salt into the flask, and add just sufficient water to moisten the salt. Through the thistle-tube, add enough concentrated sulphuric acid to cover the end of the thistle-tube.

2. Heat gently. When necessary to the maintenance of the production of gas, add more acid.

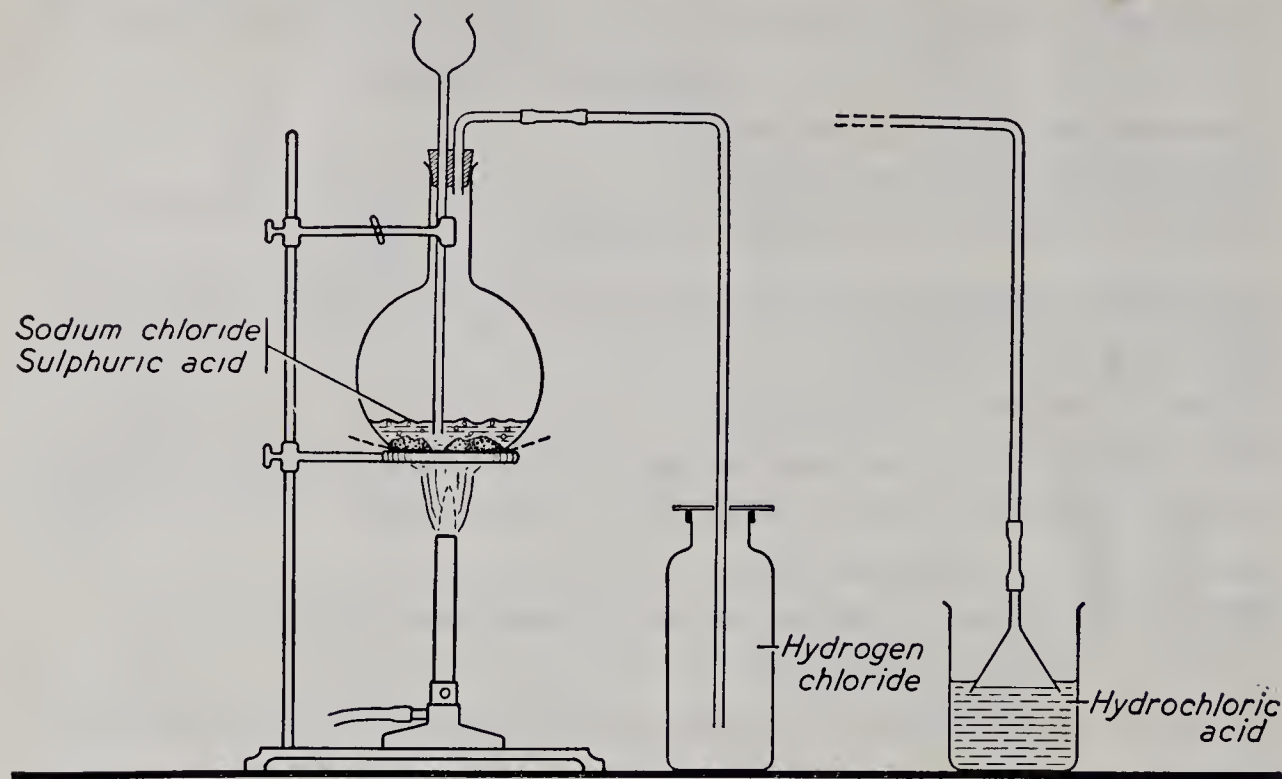


Fig. 62. The laboratory preparation and collection of hydrogen chloride.

3. Collect two bottles and two test-tubes of the gas by the upward displacement of air. Then, to obtain a solution of hydrogen chloride in water, arrange the funnel as shown in the diagram.

Describe the action in the flask when the acid is added. What is the effect of heating the mixture?

What is the colour of the gas?

What does the method of collection indicate regarding the density of the gas?

Write the equation for the laboratory preparation of hydrogen chloride.

PROPERTIES OF HYDROGEN CHLORIDE

1. Invert a test-tube of the gas in a beaker of water.

Describe the action that takes place.

What conclusion can be made regarding the solubility of hydrogen chloride in water?

2. Note the odour of the gas in the second test-tube by wafting small quantities towards the nose. Gently blow across the mouth of the tube.

Describe the odour of the gas.

Describe and explain the action that takes place when the moist breath comes in contact with hydrogen chloride.

3. Insert a blazing splint into one of the bottles of gas.

Does the gas burn?

Does the gas support combustion?

4. Moisten a piece of filter paper with ammonium hydroxide and drop the paper into the remaining bottle of gas.

Describe and account for the action that takes place.

5. Add a piece of neutral litmus paper to the solution of hydrogen chloride in water. Quarter fill four test-tubes with the solution, and test in turn with each of the following substances: magnesium, zinc, washing soda crystals, and silver nitrate solution. In the experiments with magnesium and zinc, test the gas produced with a blazing splint.

What is the effect of the solution of the gas on litmus? Name this solution.

Describe the action of magnesium and zinc on the solution. Which metal seems more active? What gas is produced? Write equations for the reactions.

Describe and account for the action that takes place with washing soda crystals. Write the equation for the reaction.

Describe and account for the action with silver nitrate solution. Write the equation for the reaction.

EXPERIMENT 81. (Demonstration) **The hydrogen chloride fountain.**

Set up the apparatus as in Fig. 63, making sure that the flask is filled with hydrogen chloride gas and that the dropper holds sufficient water. Force a few drops of water into the flask and observe the action that takes place.

Explain the formation and colour of the fountain.

QUESTIONS

1. List six properties of hydrogen chloride gas.

2. Give two identifying tests for hydrogen chloride and write chemical equations for these tests.

3. Can "fogging of the breath" be used as an identifying test for hydrogen chloride? Explain.

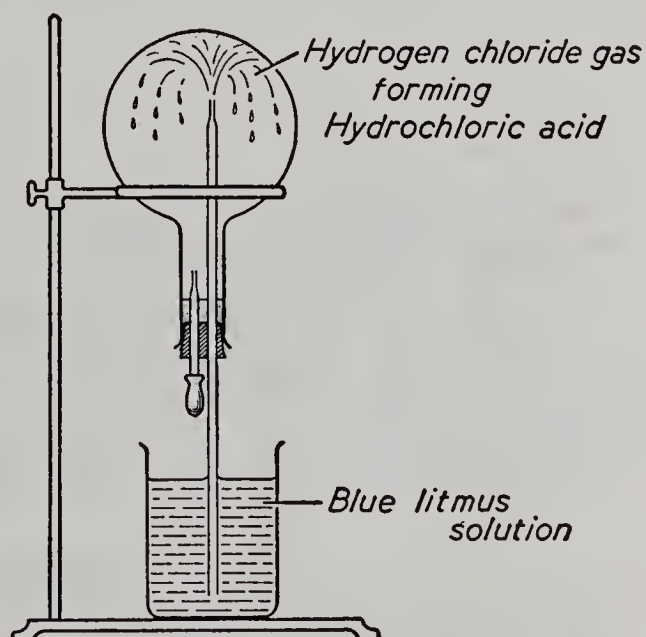


Fig. 63. The hydrogen chloride fountain.

UNIT 14

The Halogens

✓ EXPERIMENT 82. Common methods for the preparation of chlorine.

CAUTION: Chlorine is a lung-irritant gas and because of its poisonous nature should never be intentionally inhaled.

1. Place about 1 gm. of sodium chloride and 1 gm. of manganese dioxide in a test-tube and shake to mix the chemicals thoroughly. Add a few drops of concentrated sulphuric acid to the mixture. Note the colour of the gas produced and *cautiously* smell it. Test the gas with moist red and blue litmus papers. Blow your breath gently across the mouth of the test-tube.

Describe the action that takes place.

What is (a) the colour, (b) the odour of the gas?

What change takes place in the litmus paper? Did any gas, prepared previously, have this same effect on litmus?

Describe the action that takes place when the breath is blown across the mouth of the test-tube. What gas produces this effect? Is pure chlorine produced by this method, or is it mixed with another gas? Explain.

2. Place about 1 gm. of manganese dioxide in a clean dry test-tube, and add enough concentrated hydrochloric acid to cover it. Warm the mixture and test the gas as in Procedure 1.

Is the gas coloured? Is it mixed with the same gas as in Procedure 1?

3. Place a few crystals of potassium permanganate in a test-tube and add a few drops of concentrated hydrochloric acid.

Identify the gas formed in this reaction.

4. Place about 2 gm. of bleaching powder in a test-tube, add sufficient dilute sulphuric acid to cover the powder, and warm the mixture gently.

What gas is produced? How did you identify it?

QUESTIONS

1. Which of the above methods is most suitable for the laboratory preparation of chlorine? Why?

2. What is the chief objection to using Procedure 1 as a laboratory method for preparing chlorine?

3. Write chemical equations to represent each of the four methods used to prepare chlorine.

4. List three properties of chlorine as determined by this experiment.

✓ EXPERIMENT 83. (*Demonstration*) To prepare, collect, and study the properties of chlorine.

PREPARATION AND COLLECTION

Set up the apparatus as shown in Fig. 64. Put two teaspoonfuls of manganese dioxide into the flask and pour sufficient concentrated hydrochloric acid down the thistle-tube to completely cover the solid. Warm the contents of the flask and collect the gas in gas

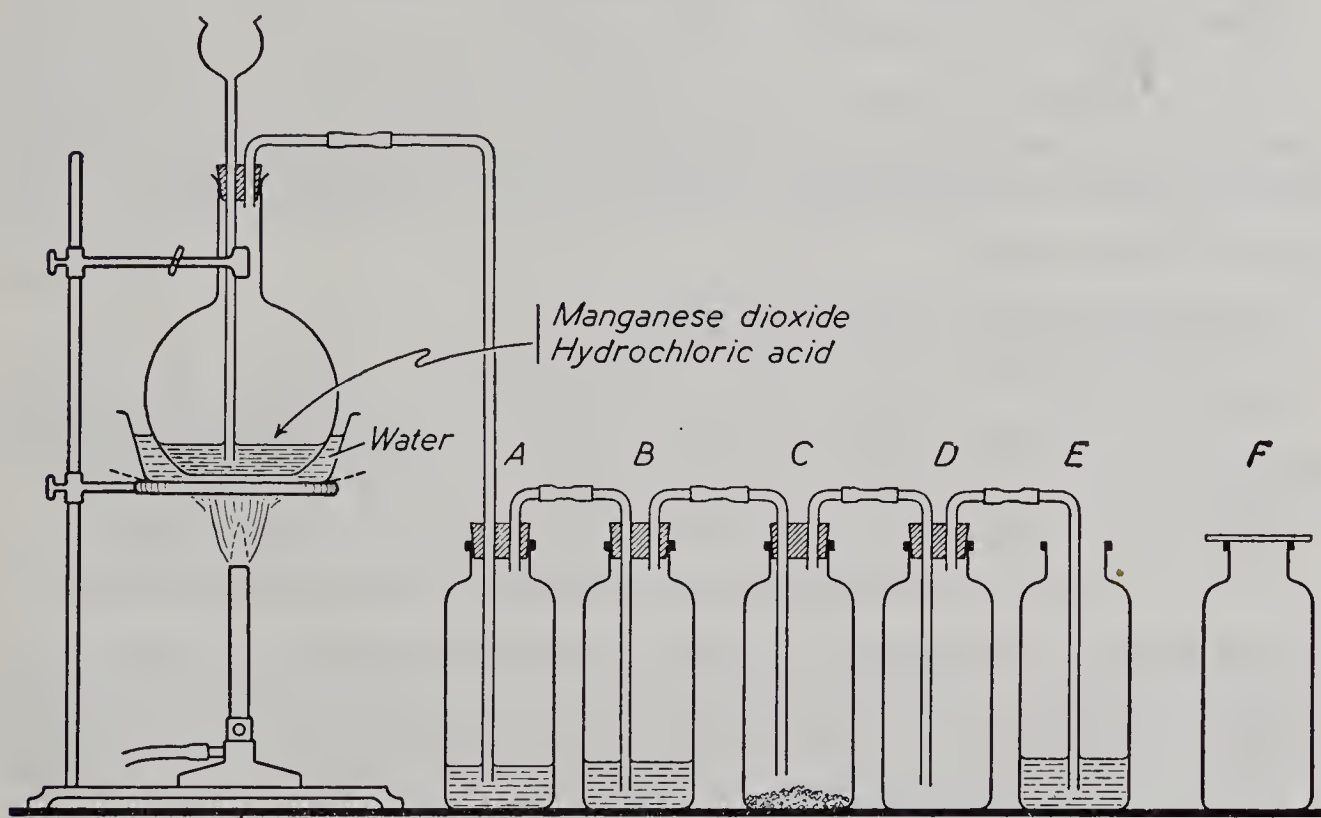


Fig. 64. The laboratory preparation and collection of chlorine. *A*, contains water to prepare chlorine water and to dissolve any hydrogen chloride that may be driven over. *B*, contains concentrated sulphuric acid to dry the chlorine. *C*, contains slaked lime for making a sample of bleaching powder. *D*, is for the collection of pure dry chlorine. *E*, contains a solution of sodium thiosulphate to prevent the escape of chlorine gas into the laboratory. *F*, a bottle of dry chlorine that has been filled at *D*.

bottles by the upward displacement of air as shown in the figure. Collect eight small bottles and one extra large bottle to be used in Procedure 6. Cover each bottle with a glass plate and allow the bottles to stand on the desk until needed.

What does the method of collection indicate regarding the density of chlorine?

What observations are noted in the flask itself?

What method is used for determining when the bottles are filled with the gas?

PROPERTIES OF CHLORINE

1. Carefully dry a small piece of sodium with filter paper to remove all traces of adhering liquid and then flatten the sodium into a thin wafer. Place this thin wafer in a deflagrating spoon and lower the spoon into a bottle of chlorine gas. If great care is used the product of this reaction may be tasted.

Describe the action that takes place.

Describe and identify the product of the reaction.

Why must great care be used in tasting the product?

Write the chemical equation for the reaction.

2. Cut a strip of copper foil one half inch wide and six inches long and after heating it in the flame of a Bunsen burner quickly drop it into a bottle of chlorine gas. After the reaction ceases add 5 ml. of water to the bottle.

Describe the action that takes place.

Describe and name the product of the reaction.

What evidence is there that copper is less active chemically than sodium?

Describe and account for the colour change when water is added.

Write the chemical equation for the action of copper on chlorine.

3. Place a small piece of yellow phosphorus about the size of a match-head in an asbestos-lined deflagrating spoon and lower it into a bottle of chlorine gas. After the reaction has ceased, add 5 ml. of water to the bottle and shake.

Describe the action that takes place.

Describe and name the product of the reaction.

Is the product of the reaction soluble in water? Explain.

Could this reaction correctly be classified as combustion? Explain.

What does this reaction indicate regarding the properties of (a) phosphorus, (b) chlorine?

Write the equation for the reaction.

4. Place a pinch of powdered antimony in the fold of a small piece of paper and slowly sprinkle the powder into a bottle of chlorine.

Describe the action that takes place.

Describe and name the product of the reaction.

What evidence is there that the reaction is exothermic?

Write the equation for the reaction.

5. (*Demonstration*) Set up a hydrogen generator and attach a glass jet of the shape shown in Fig. 65. When an inverted test-tube filled with hydrogen collected from the jet burns quietly, indicating the evolution of pure hydrogen, light the jet with the burning hydrogen and lower the burning jet into a bottle filled with chlorine. When the flame goes out remove the hydrogen jet and fill the hydrogen generator with water. Insert a piece of moist neutral litmus paper into the bottle. Blow your breath across the mouth of the gas bottle.

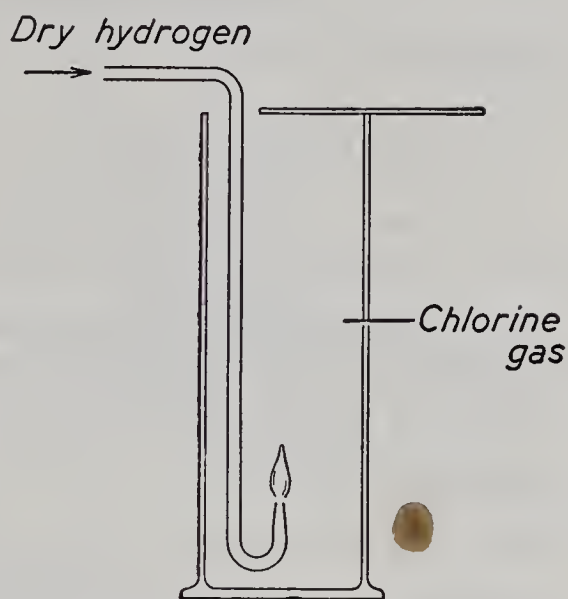


Fig. 65. The combustion of hydrogen in chlorine.

Describe the change in the appearance of the hydrogen flame when burning in chlorine.

Account for the change in colour noted in the gas bottle.

What causes the flame to be extinguished? What gas is now present in the gas bottle? Explain.

Write the chemical equation for the reaction.

6. Dip a piece of filter paper into warm turpentine, roll the paper into a loose roll sufficiently small to fit the mouth of the *large* bottle of chlorine, and quickly insert it into the bottle. After the reaction is over, blow gently across the mouth of the bottle.

Describe the action that takes place. Is it exothermic?

What gas must now be present in the bottle? Explain.

Name and account for the formation of the other product of this combustion.

What important property of chlorine is demonstrated here?

Write the chemical equation to represent this reaction.

7. Lower a burning candle or a wax taper into a bottle of chlorine and note the effect on the appearance of the flame.

Describe the change in the appearance of the flame.

Account for the reaction that takes place.

What important property of chlorine is demonstrated by this experiment?

8. Into another bottle of chlorine place a piece of lined foolscap upon which has been scribbled some pencil and ink marks. Also place in the bottle red and blue litmus papers, a cancelled postage stamp, and a piece of fruit-basket netting or other similarly dyed cloth.

Describe the effect of the dry chlorine on each of these materials.

9. Repeat Procedure 8, this time dipping each material in water before inserting it into the chlorine.

Describe the effect of the chlorine on the moistened materials.

Account for any difference in action when water is present.

Write a summation equation to explain this important chemical property of chlorine.

10. Place a piece of starch-iodized paper into a bottle of chlorine and observe the colour change in the paper. This colour change, (especially when coupled with the colour, odour, and bleaching properties) may be considered a characteristic test for chlorine.

What is the effect of chlorine on starch-iodized paper?

11. Remove the gas bottle into which slaked lime was placed (Fig. 64C), and examine the product resulting from its contact with chlorine. Place half of the material in a test-tube and add a small amount of dilute sulphuric acid. Note the odour of any gas liberated and test the gas with moist blue litmus paper.

Make a thin paste by adding water to the remaining portion of the material. Dip a piece of fruit-basket netting into the paste and then drop the cloth into dilute sulphuric acid in a beaker. Remove the cloth from the beaker, wash it in water, and observe its colour.

What is the product formed by passing chlorine over slaked lime?

Why is it an important industrial compound? How is chlorine liberated from this compound?

Write equations to represent (a) the formation of bleaching powder, and (b) the liberation of chlorine from bleaching powder.

12. To the liquid in bottle (A) (see Fig. 64) add blue and red litmus papers and a piece of coloured cotton cloth. Set the bottle aside for five minutes and examine.

What is the colour of the solution of chlorine and water?

What colour changes occur in the litmus and the cloth?

13. (*Demonstration*) Remove the chlorine generator from the system of gas bottles and bubble chlorine through water contained in a 400 ml. cylinder until the solution is saturated with chlorine. Fill a large test-tube with this solution and invert it over a funnel as shown in Fig. 66. Allow the apparatus to stand in sunlight, and after all action ceases test the gas which collects in the test-tube with a glowing splint. Test the liquid remaining in the beaker with blue litmus paper.

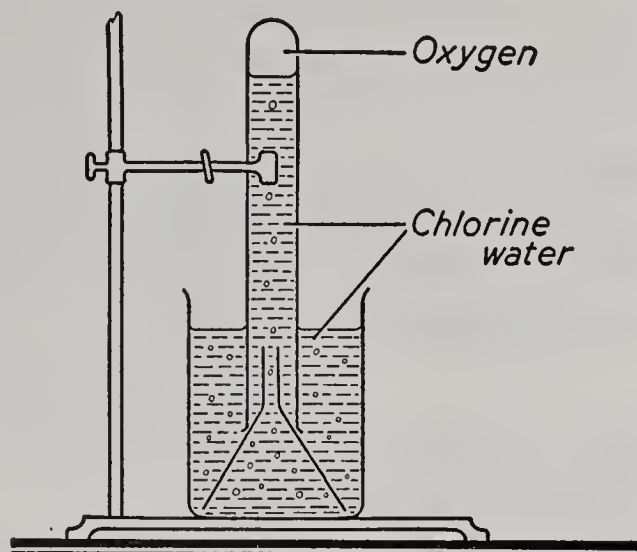


Fig. 66. The action of sunlight on a solution of chlorine and water.

Describe the action that takes place when sunlight falls on a solution of chlorine and water.

What is indicated by (a) the glowing splint test (b) the litmus test? Account for the formation of each substance.

Write equations to represent the above chemical reactions.

QUESTIONS

1. Make a diagram showing how to set up the apparatus for obtaining chlorine by the action of potassium permanganate on hydrochloric acid, and write the chemical equation for the reaction.

2. List five physical properties of chlorine.

3. List four chemical properties of chlorine.

4. What is the most important chemical property of chlorine?

5. How does bleaching with chlorine differ from bleaching with sulphur dioxide? Explain with suitable equations.

6. Explain the statement "chlorine is made by the oxidation of hydrochloric acid using manganese dioxide as the oxidizing agent".

7. What important property of chlorine is indicated by the starch-iodized paper test? Explain with the aid of a suitable equation.

✓ EXPERIMENT 84. To prepare, collect, and study the properties of bromine.

CAUTION: The vapour of bromine must not be inhaled as it is a lung-irritant. Liquid bromine causes serious flesh burns.

PREPARATION AND COLLECTION

1. Arrange the apparatus as in Fig. 67. Thoroughly mix on a piece of paper one teaspoonful of sodium bromide and one teaspoonful of manganese dioxide. Place the mixture in a large test-tube and add sufficient water to just moisten it. Add sufficient concentrated sulphuric acid to just cover the moist mixture. Quickly insert the stopper containing the delivery tube leading to the gas bottle and observe the reaction. When the reaction diminishes in activity, gently warm the test-tube.

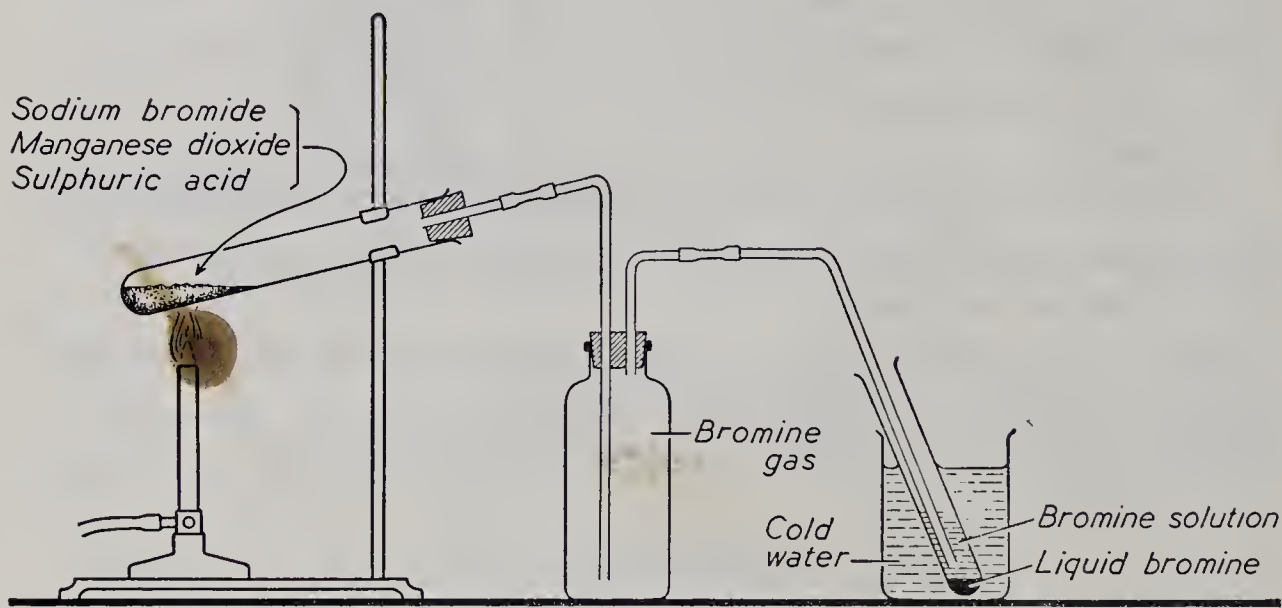


Fig. 67. The laboratory preparation and collection of bromine.

2. As the bottles become filled remove them. Collect four bottles of bromine vapour. Place a glass plate on each bottle and reserve until ready for use.

What advantage is there in adding water before the acid is added?

Describe the action that takes place within the test-tube.

Is there any evidence that hydrogen bromide is produced? Explain.

Suggest a reason for collecting the bromine as indicated in Fig. 67.

Write the chemical equation for the laboratory preparation of bromine.

PROPERTIES OF BROMINE

1. Examine the bromine in (a) the gas bottle, and (b) the test-tube.

What is the colour of bromine vapour?

What is the colour of liquid bromine?

What is the colour of bromine water?

Why is the gaseous state of bromine referred to as bromine vapour?

(TEXT p. 10)

Describe the odour of the gas.

2. Remove the glass plate from a bottle of bromine vapour. Place moist blue and red litmus papers in the bottle and allow it to stand.

Describe and explain the action on the litmus papers.

NOTE: For the following experiments it may be more convenient to use a laboratory supply of bromine.

3. To a bottle of bromine vapour add water to a depth of one inch and shake the bottle vigorously. Drop a piece of blue litmus paper and a piece of red litmus paper into the solution.

What evidence is there to show that bromine is soluble in water?

Explain the changes in colour of the litmus papers.

4. To a bottle of bromine vapour add 5 ml. of carbon tetrachloride and shake the bottle vigorously.

Is bromine soluble in carbon tetrachloride?

How does the action compare with that of water?

5. Sprinkle a pinch of powdered antimony into a bottle of bromine vapour.

Describe the action that takes place.

How does this action compare with the action of antimony on chlorine?

6. Place a small piece of yellow phosphorus about the size of a match-head in an asbestos-lined deflagrating spoon and lower the spoon into a bottle of bromine vapour.

Describe the action that takes place.

What is the nature and the name of the product?

7. Carefully remove the test-tube containing the liquid bromine from the beaker. Note the colour of the bromine solution above the liquid bromine. Add a few drops of bromine to a small portion of starch "solution".

Why is caution necessary in removing liquid bromine from a container?

Describe the appearance of the starch "solution" before and after the addition of bromine.

8. Place 3 ml. of bromine water into a test-tube and drop a piece of starch-iodized paper into the solution.

Describe and account for the colour change in the starch-iodized paper.

Can this colour change be used as a distinctive test for bromine as it is for chlorine? Explain.

What does this action of bromine indicate regarding the relative activities of bromine and iodine? Explain with the aid of a chemical equation.

9. Into a test-tube half-filled with water place a level teaspoonful of sodium bromide, and shake. To the resulting solution add 2 ml. of carbon tetrachloride. To the liquids in the test-tube add sufficient freshly prepared chlorine water to three-quarters fill the test-tube, and shake vigorously.

What is the colour of sodium bromide solution?

What is the colour of carbon tetrachloride after it is placed in the sodium bromide solution?

What happens when the chlorine water is added? Explain.

What happens when the test-tube is shaken? Explain.

What does this experiment indicate about the relative activities of chlorine and bromine? Explain with the aid of a chemical equation.

QUESTIONS

1. List the physical and chemical properties of bromine.
2. What experimental evidence is there to indicate that bromine is less active chemically than chlorine?

✓ EXPERIMENT 85. To prepare, collect, and study the properties of iodine.

PREPARATION AND COLLECTION

Place a teaspoonful of potassium iodide and a teaspoonful of manganese dioxide in a clean dry 250-ml. beaker and shake the

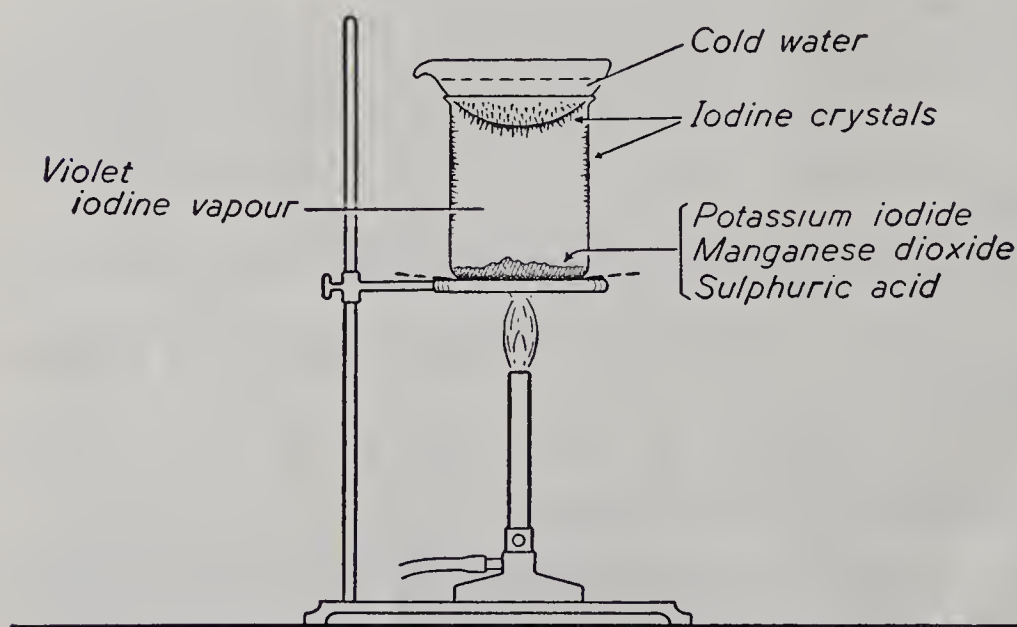


Fig. 68. The laboratory preparation and collection of iodine.

beaker to mix the two chemicals. Nearly fill an evaporating dish with cold water and arrange the beaker and evaporating dish as in

Fig. 68. Lift the evaporating dish from the beaker and add approximately 10 ml. of concentrated sulphuric acid to the mixture in the beaker. Carefully replace the evaporating dish on top of the beaker and apply very gentle heat to the beaker for a short time. Then remove the flame.

Describe the action that takes place inside the beaker.

What is the colour of iodine vapour?

What is the colour of iodine crystals? Where are the crystals found? Account for their formation.

Write the chemical equation for the laboratory preparation of iodine.

PROPERTIES OF IODINE

1. Place a few crystals of iodine in a clean dry test-tube. Hold the test-tube at a 45° angle and gently heat.

Describe all the changes that take place in the iodine and account for these changes.

Why is the iodine heated gently? Why is the test-tube held at an angle?

Describe the crystal structure of iodine.

2. Place a few crystals of iodine in each of seven clean dry test-tubes. Then using one test-tube at a time, test the solubility of iodine in each of the following liquids: water, a solution of potassium iodide, ethyl alcohol, ether, chloroform, carbon tetrachloride, and carbon disulphide. Retain each of the solutions for Procedure 5.

Describe the visible changes in each case.

Which solvent seems to be (a) the best, (b) the poorest?

Could solubility in any of these solvents be used as an identifying test for iodine? Explain.

3. Place a few crystals of iodine in a mortar along with a small drop of mercury and grind the two together with a pestle.

CAUTION: The product of this reaction is a poison.

Describe and interpret the result.

Write the chemical equation for the reaction.

4. (*Demonstration*) Place a piece of yellow phosphorus about the size of a grain of wheat in a clean dry evaporating dish and carefully sprinkle a few crystals of iodine on the phosphorus.

Describe and interpret the result.

Write the chemical equation for the reaction.

5. Put a teaspoonful of corn starch into a 250-ml. flask and half fill the flask with water. Heat the starch-water mixture until the mixture appears translucent. Add sufficient water to form a mixture that pours easily. Test separate quantities of this starch paste or "solution" with each of the seven solutions retained from Procedure 2. If necessary, dilute the contents of each test-tube with water until a recognizable colour is produced.

Account for the change in appearance of the starch when heated with water.

Describe the colour effect in each test-tube.

What is the effect of water dilution on the colour?

Account for the starch-iodized paper test used with chlorine and bromine.

How would you test for the presence of starch in a food sample?

QUESTIONS

1. What procedure would be followed in purifying a sample of iodine?

2. How would you distinguish between solutions containing chlorine, bromine, and iodine?

3. Compare the properties of chlorine, bromine, and iodine, showing how the properties change with increasing atomic weights. (TEXT p. 236)

✓ EXPERIMENT 86. Tests for chlorides, bromides, and iodides.

1. Into each of three separate test-tubes put (a) sodium chloride, (b) sodium bromide, and (c) sodium iodide, to a depth of $\frac{1}{4}$ inch, and to each test-tube add sufficient distilled water to three-quarters fill the tube. Divide each solution into three parts and arrange three sets of solutions so that each set contains solutions of the three salts. Use one set of solutions in each of the following procedures.

2. Bubble chlorine (or add freshly prepared chlorine water) into each solution in turn. After any visible changes have taken place add a small amount of carbon tetrachloride to each test-tube.

Describe the visible change that takes place in each test-tube.

What inference can be drawn from each observation?

What conclusion can be drawn from the effect produced upon the addition of carbon tetrachloride? Account for these conclusions.

Write chemical equations to express the reactions that have taken place.

Is this a good test for a chloride?—a bromide?—an iodide? Explain.

3. To the second set of solutions add small amounts of carbon tetrachloride and observe. To each test-tube add 2 ml. of bromine water, shake and observe.

What effect has carbon tetrachloride on a solution of (a) sodium chloride, (b) sodium bromide, and (c) sodium iodide?

What is the result of the addition of the bromine water?

What conclusion could be drawn from this experiment?

Could this test be used for an iodide? Write an equation to explain your answer.

4. To each test-tube of the third set of solutions, add 1 ml. of dilute nitric acid followed by 5 ml. of silver nitrate solution. Observe each reaction closely. Continue to watch the reactions, holding the test-tubes in sunlight if possible. Retain the test-tubes for Procedure 5.

Give all visible changes taking place in each test-tube.

Could the visible changes be used to identify separate samples of chlorides, bromides, and iodides? Explain.

Write chemical equations to express each of the reactions.

What is the effect of holding the test-tubes in sunlight?

Why is it necessary to add nitric acid? (TEXT p. 232)

5. To each test-tube add 5 ml. of ammonium hydroxide solution and shake each test-tube to mix its contents thoroughly.

What is the effect of the addition of ammonium hydroxide to each test-tube? Explain.

Describe how the addition of ammonium hydroxide may be considered as a confirmatory test.

QUESTIONS

1. What is the best test for a chloride?
2. What is the best test for a bromide?
3. What is the best test for an iodide?

EXPERIMENT 87. (*Demonstration*) To etch glass with hydrogen fluoride.

CAUTION: Hydrogen fluoride and its aqueous solution hydrofluoric acid are both very corrosive and must neither be inhaled nor allowed to come in contact with the skin.

1. Coat the inside of an evaporating dish with a generous layer of paraffin. Select a piece of window glass large enough to cover the

evaporating dish adequately, and place an even coating of paraffin over its surface. With a pin or a fine stylus, scratch a design through the wax to the glass plate.

2. Place approximately 3 grams of calcium fluoride in the paraffin-lined dish and add sufficient concentrated sulphuric acid to make a thin paste.

3. Gently press the glass plate, wax-side down, on the evaporating dish and set it in a fume cabinet or beside an exhaust vent for a time. After about an hour, scrape or melt the paraffin from the plate and examine the surface of the glass.

Describe the appearance of the plate and account for the change that has taken place.

Why is it important to line the evaporating dish with paraffin?

How should hydrofluoric acid be stored in the laboratory?

UNIT 15

Compounds of Nitrogen

✓ EXPERIMENT 88. To prepare, collect, and study the properties of ammonia.

PREPARATION AND COLLECTION

1. Place as much slaked lime as will lie on a ten-cent piece in the palm of one hand. In the other hand, place an equal quantity of ammonium chloride. Holding the hands over the sink, rub the two chemicals together vigorously, and then **cautiously** note the odour of the gas produced.

What is the chemical name for slaked lime?

Describe the odour of the gas produced. Name the gas.

What three purposes are served by rubbing the hands vigorously?

Write the equation to represent the reaction.

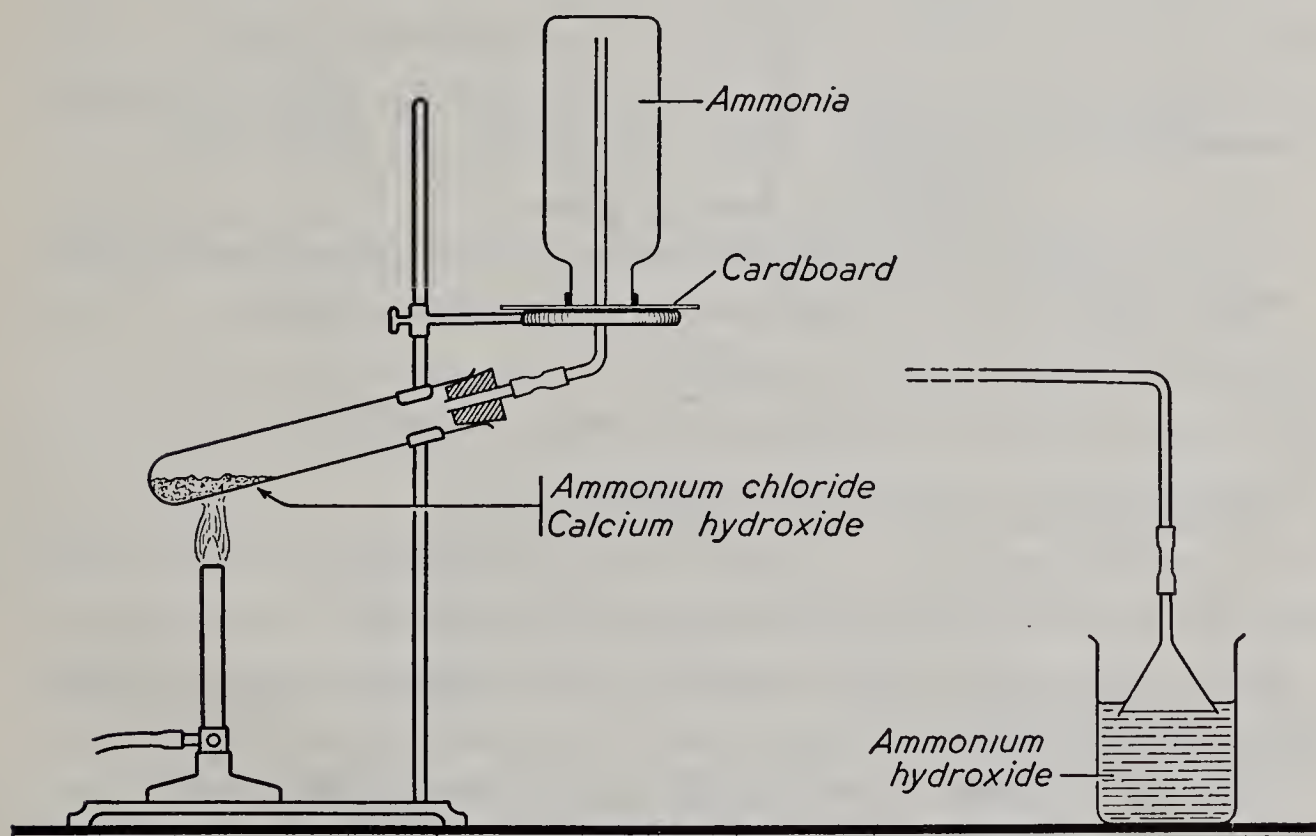


Fig. 69. The laboratory preparation and collection of ammonia.

2. Quarter fill a Pyrex test-tube with ammonium chloride, add an equal quantity of calcium hydroxide, and shake the test-tube so as to

mix the two solids thoroughly. Set up the apparatus as shown in Fig. 69; heat the mixture gently, and collect three bottles of ammonia by the downward displacement of air. To determine when a bottle is filled, hold a piece of moist red litmus paper at the mouth of the bottle. When the bottles are filled with ammonia, place them mouth down on the table. Then, as illustrated in the diagram, attach the delivery tube containing the funnel to the generating test-tube, and immerse the mouth of the funnel in water in a beaker.

What is the effect of ammonia on moist red litmus?

PROPERTIES OF AMMONIA

1. Hold a bottle of ammonia mouth down. Insert a blazing splint upward into the bottle and observe.

What is the colour of ammonia?

What two properties of ammonia are illustrated by using the blazing splint?

Why is the bottle held mouth down?

Distinguish between "ammonia" and "ammonium".

2. Put four drops of concentrated hydrochloric acid into a gas bottle, cover with a glass plate, and shake the bottle. Place a bottle of ammonia mouth up on the table. Invert the bottle containing the acid over the ammonia bottle and remove the glass plate.

What gas is given off by the concentrated acid?

Describe and account for the action that takes place within the bottles.

Which bottle shows the most pronounced change? Why?

Write the equation for the reaction.

3. Invert a bottle of ammonia in a beaker of water.

Describe and account for the change observed.

4. Test the solution of ammonium hydroxide obtained in the preparation and collection of ammonia with neutral litmus paper.

Write the equation that represents the reaction when ammonia "dissolves" in water.

What is the colour, odour, and effect on litmus of this solution?

5. Add 10 ml. of a solution of ammonium hydroxide to 10 ml. of a dilute solution of hydrogen chloride.

Write the equation for the reaction that takes place. Why does the ammonium chloride that is formed not appear as a white solid as it did in Procedure 2?

6. (*Demonstration*) Saturate a pile of a dozen pieces of filter paper with a concentrated solution of ammonium hydroxide, and place the pile of paper on a glass plate at one end of the table. At the opposite end of the table, saturate a pile of a dozen pieces of filter paper with concentrated hydrochloric acid. Now bring the two piles of filter paper, together, one in each hand, and press them firmly against each other. Observe the reaction that takes place.

Describe and account for what happens.

What evidence is there that the reaction is exothermic?

EXPERIMENT 89. (*Demonstration*) The ammonia fountain.

Set up the apparatus as in Fig. 70, making sure that the flask is filled with ammonia gas and that the dropper holds sufficient water. Force a few drops of water into the flask and observe the action that takes place.

Explain the formation and colour of the fountain.

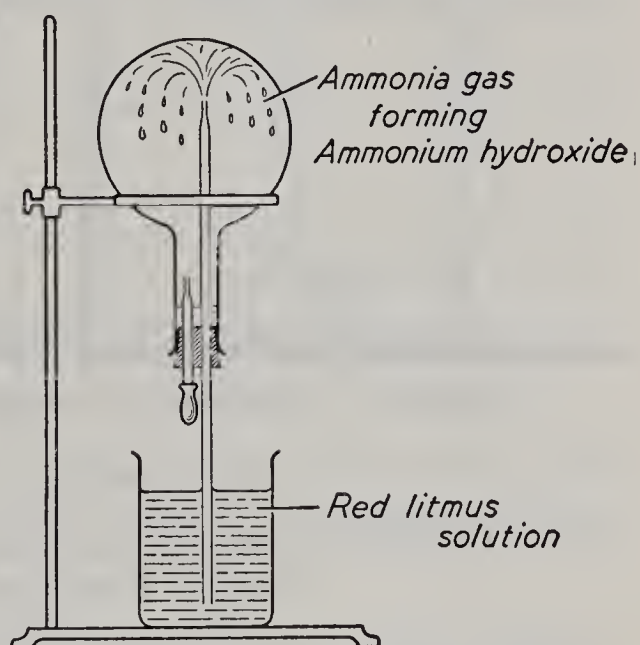


Fig. 70. The ammonia fountain.

QUESTIONS

1. Why is the heating of ammonium hydroxide not considered a laboratory method of preparing ammonia even though it is a good source of ammonia?

2. Write nine equations to represent the production of ammonia gas, using three different ammonium salts and three different bases.

✓ EXPERIMENT 90. To prepare, collect, and study the properties of nitric acid.

PREPARATION AND COLLECTION

1. Arrange the apparatus as shown in Fig. 71. Place two teaspoonfuls of sodium nitrate in the retort, and add sufficient concentrated sulphuric acid to cover the sodium nitrate.

2. Heat the mixture in the retort with a moderately hot flame, and observe the action in the retort and in the condensing flask.

3. When no more nitric acid collects in the flask, remove the flame, and examine the contents of the retort and of the condensing flask.

Describe the action that takes place in the retort.

What evidence is there that nitric acid has a low boiling point?

Describe the product remaining in the retort after the reaction.

Describe the nitric acid in the condensing flask.

Write the equation for the preparation of nitric acid.

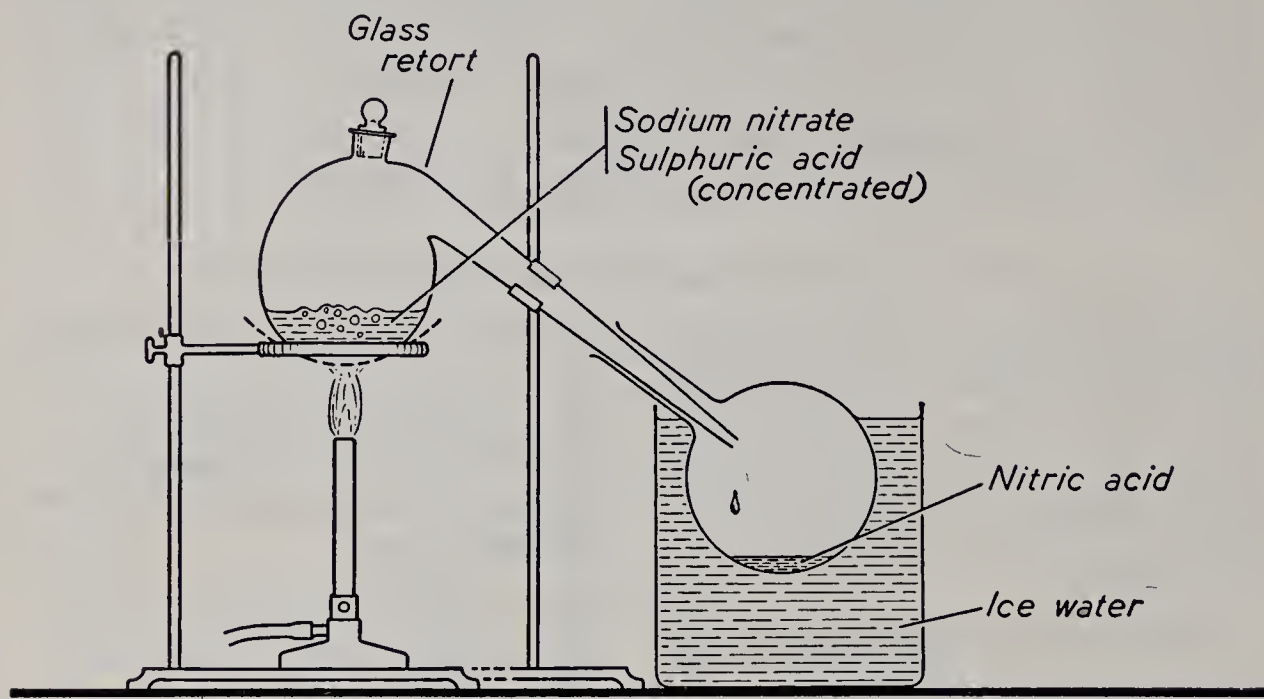


Fig. 71. The laboratory preparation and collection of nitric acid.

PROPERTIES OF NITRIC ACID

CAUTION: Concentrated nitric acid is very corrosive and must be handled with great care.

Because the amount of nitric acid prepared in this experiment is small, it is more convenient to use concentrated nitric acid from the reagent bottle for most of the following experiments.

1. Place some white of egg in an evaporating dish and add 2 ml. of concentrated nitric acid to the egg white. Heat the evaporating dish and observe the changes that take place.

Wash the excess nitric acid from the egg white and then add 2 ml. of ammonium hydroxide and observe any further changes.

Describe the effect of nitric acid on egg white.

What happens when ammonium hydroxide is added?

Egg white is a protein and most proteins react in this way with concentrated nitric acid. The reaction is known as the Xanthoproteic reaction.

2. Add one drop of concentrated nitric acid to one-third of a test-tube of water. Taste a drop of this solution. Note the effect of this solution on neutral litmus paper.

Describe the taste of the nitric acid solution.

What is the effect on litmus?

3. (*Demonstration*) **CAUTION:** When nitric acid boils, the liquid may spatter from the test-tube.

Heat 2 ml. of concentrated nitric acid in a Pyrex test-tube until it boils, and observe. Hold a glowing piece of charcoal, which has been formed on the end of a long charred wooden splint, in the vapour above the boiling acid.

Describe all the changes that take place. Account for any colour change.

Write the equation for the decomposition of nitric acid by heat.

Account for the action of the glowing charcoal. Write the equation that represents the reaction.

What is the effect of the nitric acid vapour on the part of the wood that is not charred?

4. Drop a small piece of zinc into 2 ml. of dilute nitric acid in a test-tube, observe the reaction, and test the escaping gas with a blazing splint.

Describe the reaction in the test-tube. Is hydrogen produced?

5. Place a small piece of copper in each of two separate test-tubes. To one test-tube add dilute nitric acid, and to the other, concentrated nitric acid, and observe.

What is observed in each test-tube?

Name the products of each reaction, and write the equations.

6. (*Demonstration*) Test a few drops of concentrated nitric acid with a solution of barium chloride to find out if it contains any trace of a sulphate. To 2 ml. of the same acid, add a piece of roll sulphur about the size of a bean, and boil the acid for several minutes. Test the resulting solution with barium chloride solution. Compare the resulting precipitate with the precipitate formed (if any) when the nitric acid was initially tested.

Describe and account for any reaction between barium chloride and the nitric acid tested. Why does commercial nitric acid often contain traces of a sulphate?

Describe and account for the observations noted when barium chloride solution is added after the nitric acid has been boiled with sulphur.

What important property of nitric acid is shown by this experiment?

Write the equation for the action of concentrated nitric acid on sulphur.

QUESTIONS

1. Why must nitric acid be handled with extreme care?
2. Why is sulphuric acid rather than any other acid used in the preparation of nitric acid?
3. Why is the apparatus used in the preparation of nitric acid made entirely of glass?
4. Explain why hydrogen is not produced when metals are added to either dilute or concentrated nitric acid. (TEXT p. 252)

✓ EXPERIMENT 91. The properties of nitrates.

1. Examine samples of each of the following nitrates: sodium nitrate, potassium nitrate, lead nitrate, copper nitrate, and barium nitrate.

Describe each of the nitrates selected.

2. Quarter fill separate Pyrex test-tubes with each nitrate, and heat each intensely. Hold a glowing splint to the mouth of each test-tube in turn.

Describe and account for the action of the glowing splint.

Write a chemical equation for each reaction.

Describe and account for the action in each case.

3. Add a few crystals (or a solution) of sodium nitrate to 5 ml. of

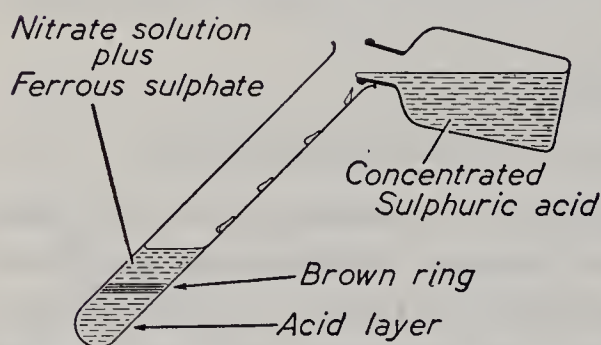


Fig. 72. Testing for nitrates (nitrate ion).

a concentrated, freshly-prepared solution of ferrous sulphate in a test-tube, and shake to mix thoroughly. Clamp the test-tube in an inclined position of about 60° to the vertical, and slowly pour 2 ml. of concentrated sulphuric acid down the side of the tube so that the heavy acid will underlie the solution without much mixing.

Describe the appearance at the junction of the acid and the solution.

Why is sulphuric acid added to the mixture in testing for a nitrate?

This procedure constitutes a delicate test for the nitrate ion.

✓ EXPERIMENT 92. (Demonstration) To prepare, collect, and study the properties of nitrous oxide.

PREPARATION AND COLLECTION

CAUTION: Ammonium nitrate may explode if heated too rapidly.

1. Place about 15 gm. of ammonium nitrate in a Pyrex test-tube and arrange the apparatus as shown in Fig. 73. Heat the ammonium nitrate gently until it melts, being very careful not to overheat.

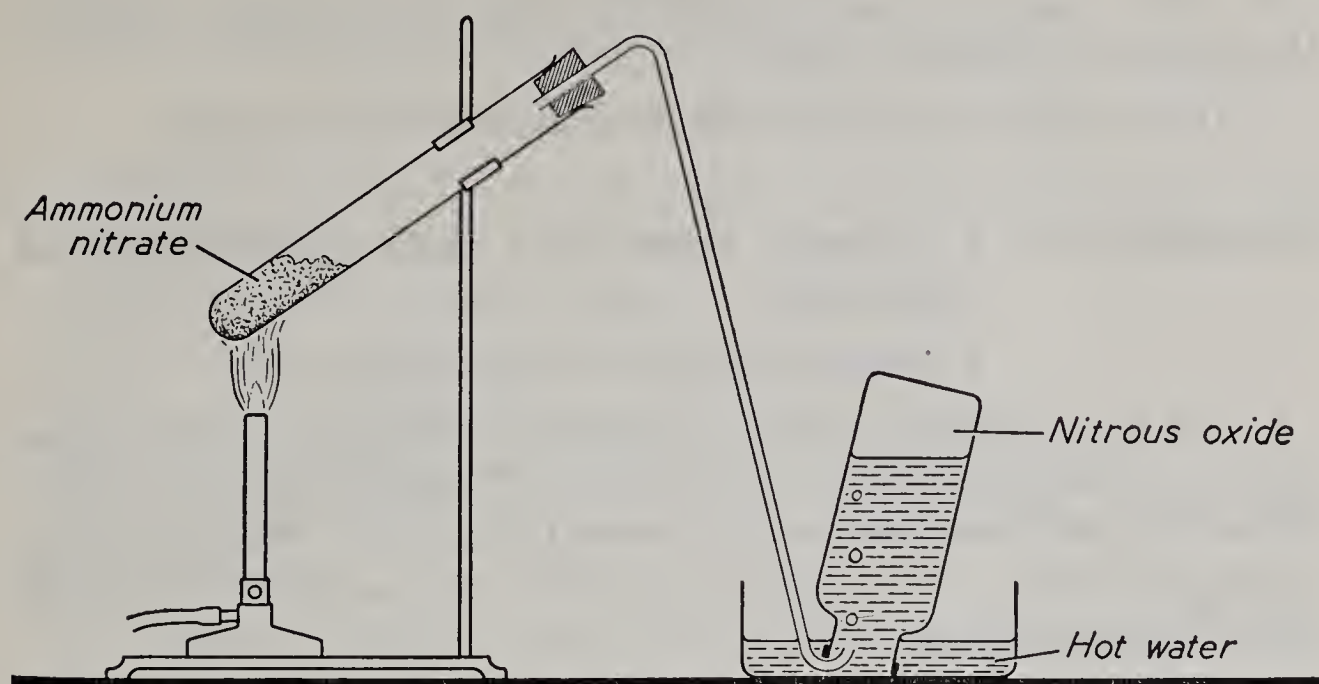


Fig. 73. The laboratory preparation and collection of nitrous oxide (laughing gas).

2. Collect three bottles of the gas by the downward displacement of hot water. Place a glass plate over the mouth of each bottle and set the bottles upright on the table.

PROPERTIES OF NITROUS OXIDE

1. Observe the colour and odour of nitrous oxide and invert a bottle of the gas in a dish of cold water.

Describe the colour and odour of nitrous oxide.

Account for the action that takes place when a bottle full of the gas is inverted in cold water.

Why is hot water used in the original collection?

2. Insert a glowing splint into a bottle of the gas and observe.

Does nitrous oxide support combustion?

Does nitrous oxide burn?

3. Place a piece of yellow phosphorus, the size of a grain of wheat, on an asbestos-lined deflagrating spoon. Ignite the phosphorus and lower the burning phosphorus into a bottle of nitrous oxide.

Describe and account for the action that takes place.

QUESTIONS

1. Write the equation for the laboratory preparation of nitrous oxide.
2. Write the equation for the burning of (a) carbon, and (b) phosphorus, in nitrous oxide.
3. How could nitrous oxide be distinguished from oxygen?

✓ **EXPERIMENT 93. To prepare, collect, and study the properties of nitric oxide.**

PREPARATION AND COLLECTION

1. Put approximately 20 gm. of copper turnings in a 250-ml. flask and set up the apparatus as in Fig. 74. Through the thistle-tube add sufficient dilute nitric acid to cover the copper. The flask may be warmed gently, but if allowed to stand, the reaction will proceed with increasing speed.

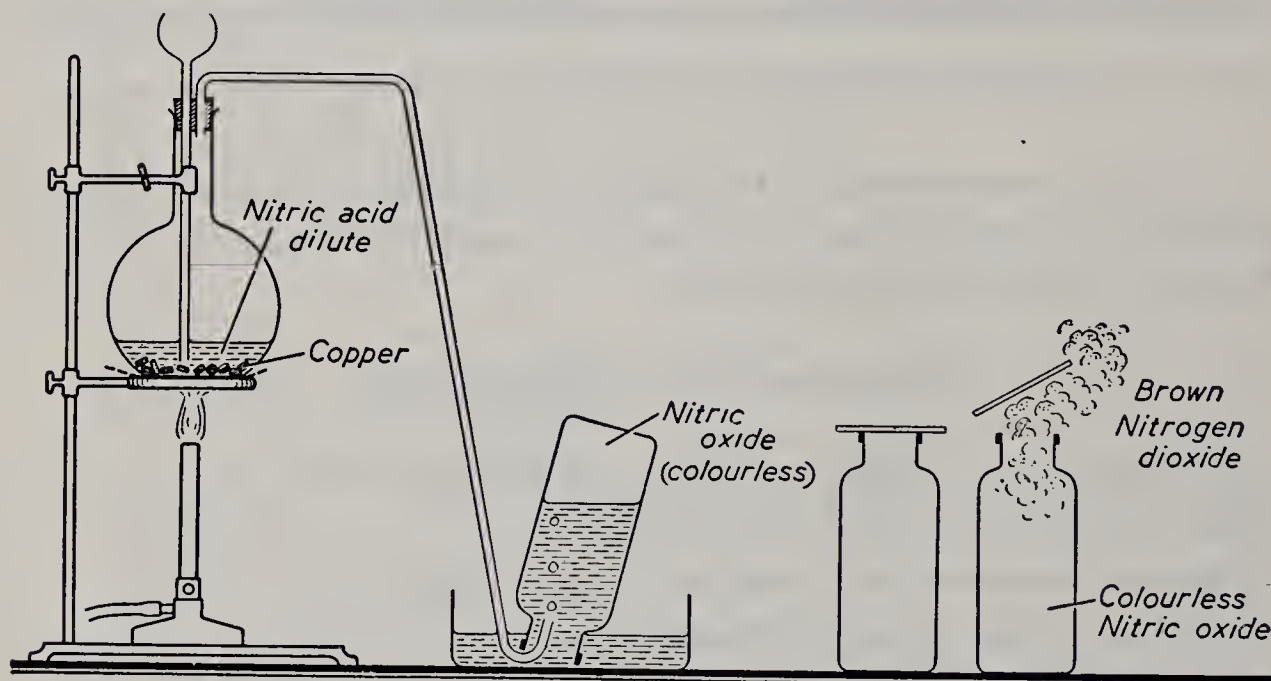


Fig. 74. The laboratory preparation and collection of nitric oxide.

2. Observe the changes that occur inside the flask. Collect four bottles of the gas by the downward displacement of water. Cover three of the bottles with glass plates and place them upright on the table. Allow the fourth bottle to remain in the pneumatic trough until ready for use.

Describe the action of copper on dilute nitric acid.

What colour are the bubbles of gas that form in the liquid?

What is the colour of the solution?

What indication is there that the reaction is exothermic?

Describe and account for the appearance of the gas inside the flask.

What is the colour of the gas collected in the gas bottles? What has happened to the coloured gas inside the flask?

Write the equation for the preparation of nitric oxide.

PROPERTIES OF NITRIC OXIDE

1. Observe the colour of nitric oxide and remove the glass plate from one of the bottles.

What is the colour of nitric oxide?

What happens when the glass plate is removed?

Why is it impossible to note the odour of nitric oxide?

Name the gas produced when nitric oxide is exposed to air.

Write the equation for the reaction that takes place when nitric oxide is exposed to air.

2. Lower (a) a blazing splint, and (b) burning yellow phosphorus in an asbestos-lined deflagrating spoon, into separate bottles of nitric oxide.

Does the gas burn?

Does nitric oxide support the combustion of (a) wood, (b) phosphorus?

3. Allow a few bubbles of oxygen from an oxygen generator to enter the bottle containing the nitric oxide remaining in the pneumatic trough. Observe the change that takes place and continue to add oxygen, a few bubbles at a time, until no further change occurs.

Describe and account for the change that takes place.

Write the equation for the reaction when the two gases mix.

QUESTIONS

1. How does the decomposition temperature of nitric oxide compare with that of nitrous oxide? Explain your answer.

2. How could nitric oxide be used as a test for oxygen? Explain with the aid of an equation.

EXPERIMENT 94. The preparation of nitrogen dioxide.

To a few copper turnings in a test-tube, add 2 ml. of concentrated nitric acid, and observe the action that takes place. Lower a blazing splint into the test-tube.

What is the colour of the bubbles of gas that appear in the nitric acid? What is the colour of the solution?

What change takes place in the colour of the solution as the reaction continues? Explain.

Does nitrogen dioxide burn? Does it support combustion?

Is nitrogen dioxide more dense or less dense than air?

What method could be used to collect a bottle of the gas? Why?

Write the equation for the laboratory preparation of nitrogen dioxide.

UNIT 16

Calcium and Magnesium

EXPERIMENT 95. To prepare a sample of calcium carbonate (precipitated chalk).

1. Mix together 50 ml. of a saturated solution of calcium chloride, and 50 ml. of a saturated solution of sodium carbonate in a 250-ml. beaker, stir, and observe.

2. Filter the mixture. Rinse the beaker with water to make sure that all of the contents of the beaker are transferred to the filter paper.

3. Wash the residue on the filter paper several times with water and allow the residue on the filter paper to dry thoroughly. Scrape the residue onto a watch-glass and note its appearance.

4. To a small portion of the residue in a test-tube add 2 ml. of dilute hydrochloric acid. Hold a drop of limewater on a glass tube in the mouth of the test-tube and observe.

Describe what happens when the two solutions are mixed.

What is the colour of the precipitate?

Why was the residue washed with water?

Describe the appearance of the dried residue.

Describe and account for the action of the acid on the residue.

Write the chemical equation to represent this method of making calcium carbonate.

What property of calcium carbonate makes possible its manufacture by this method?

EXPERIMENT 96. To prepare quicklime from calcium carbonate.

1. Moisten a piece of marble (calcium carbonate) about the size of a bean, and touch it to a piece of red litmus paper.

2. Put the piece of marble into a clean crucible and heat it intensely for twenty minutes.

3. Allow the crucible and contents to cool to room temperature and examine the substance in the crucible.

4. Touch one edge of the material in the crucible with a piece of moist red litmus paper and examine the litmus paper.

5. Allow a drop of water to fall on the material in the crucible and observe. Add a few more drops of water and note if there is any further change. Then add 5 ml. of water, stir, and filter the mixture. Catch the filtrate in a test-tube. Blow your breath through the filtrate and observe.

Does the marble chip have any effect on moist litmus?

Describe the changes observed in the marble when it is strongly heated.

Name and describe the product remaining in the crucible after it cools.

Does the product remaining in the crucible have any effect on moist red litmus?

Write the equation for the preparation of quicklime.

Describe the action that takes place when a drop of water is added to the substance in the crucible.

What is observed when the breath is blown through the filtrate? Name the filtrate.

EXPERIMENT 97. The slaking of lime.

CAUTION: When lime is slaked a minor explosion may occur.

Place a piece of quicklime the size of a walnut in an evaporating dish. Protecting your eyes with a glass plate, drop 4 drops of water on the lime and observe. Continue to add water a few drops at a time as long as any change takes place. Finally add about 20 ml. of water, stir the mixture, and filter some of it into a test-tube. Blow your breath through the filtrate.

Describe all the changes that take place in the quicklime when water is slowly added.

What evidence is there that the slaking of lime is an exothermic reaction?

Describe the slaked lime prepared and compare it with the quicklime used in its preparation.

Write the equation for the slaking of lime.

What is the name of a water solution of slaked lime?

What two changes must take place before quicklime can be changed to limewater?

EXPERIMENT 98. The making of bleaching powder.

1. One-third fill a test-tube with calcium hydroxide. Support the test-tube in a horizontal position and spread the calcium hydroxide evenly along the tube.

2. Place a glass tube leading from a chlorine generator into the test-tube so that the end of the tubing is near the closed end of the test-tube. Slowly pass chlorine into the test-tube for at least ten minutes.

3. Empty the freshly made bleaching powder from the test-tube into a beaker and note its odour. Add water so as to make a thin paste and divide the paste into two parts.

4. To one part of the mixture add 5 ml. of dilute sulphuric acid. Cautiously note the odour of the gas produced and hold a piece of moist blue litmus paper in the gas.

5. Place a piece of fruit-basket netting, or some other coloured cloth, in the second portion of the paste, and after it is well coated with paste, drop it into a beaker containing dilute sulphuric acid. Remove the cloth, wash it, and note any colour change.

Is there any noticeable change in the slaked lime when chlorine is passed over it?

Describe and account for the effect of sulphuric acid on bleaching powder.

Explain the action on the moist litmus paper.

Describe and account for the colour changes in the cloth.

Write the equation for the preparation of bleaching powder from slaked lime and chlorine.

Write the equation for the liberation of chlorine from bleaching powder.

EXPERIMENT 99. The setting of plaster of Paris.

1. Examine a lump of gypsum, and a sample of plaster of Paris.

From a comparison of physical properties, how would you distinguish gypsum from plaster of Paris?

2. Place two teaspoonfuls of plaster of Paris in an evaporating dish and add water, a few drops at a time, stirring the mixture constantly. When the mixture is in the form of a thick paste, transfer it to a glass plate which has been lightly coated with vaseline (petroleum jelly).

3. Press into the plaster of Paris paste a coin (or other object) which has been lightly coated with vaseline. Scrape away any projections of plaster of Paris that might interfere with the removal of the coin when the plaster of Paris hardens. After the plaster of Paris has hardened or "set", carefully pry the coin from its position and observe the imprint of the coin.

What useful purpose is served by the vaseline?

Why is the imprint of the coin perfect in every detail?

What causes the plaster of Paris to harden? Write the equation that represents the setting of plaster of Paris.

EXPERIMENT 100. Water of temporary hardness.

1. Quarter fill a test-tube with distilled water and add one drop of soap solution. Cover the mouth of the tube with your thumb and shake the tube vigorously back and forth ten times. Observe the effect of the soap on the distilled water. Continue to add soap solution, one drop at a time, shaking after each addition. Count the number of drops of soap solution required to produce a permanent lather. A permanent lather is one that will last for at least one minute.

What effect is produced when soap solution is added to distilled water?

How many drops of soap solution are required to produce a lather that lasts for at least one minute?

Why is distilled water considered to be "soft"? (TEXT p. 264)

2. Half fill a large test-tube with limewater and slowly bubble carbon dioxide through the limewater until the white precipitate that forms, dissolves. The water now contains a solution of calcium bicarbonate and is said to have "temporary hardness". Divide the clear solution of calcium bicarbonate into four equal parts by pouring it into four separate test-tubes.

Write equations to represent (a) the formation of the white precipitate, and (b) the "dissolving" of the precipitate.

3. To one of the test-tubes from Procedure 2, add soap solution one drop at a time. Count the drops as they are added and after adding each drop, hold the thumb over the mouth of the test-tube and shake the tube vigorously ten times. Continue adding the soap solution until a lather that lasts for at least one minute is produced. Record the number of drops of soap required to produce this permanent lather.

Describe the action of the soap on the hard water.

When does the soap produce a lather with the water? How many drops of soap solution were required to produce this lather?

Explain how the addition of soap softens water of temporary hardness.

4. Hold a second test-tube containing water of temporary hardness in the flame of a Bunsen burner and slowly heat until it comes to the

boiling point. Observe the changes that take place and cool the solution to room temperature. Add soap solution to the test-tube as in Procedure 3, and count the number of drops required to produce a permanent lather.

Describe and account for the changes that take place when water of temporary hardness is boiled. Write the equation to represent the reaction.

How many drops of soap solution were required to produce a permanent lather?

5. To one of the remaining test-tubes, add a crystal of washing soda and shake the test-tube. Observe any changes in the contents of the test-tube and test with a soap solution as in Procedure 3.

Describe and account for any changes that take place.

Write the equation for the action of washing soda on water of temporary hardness.

How many drops of soap solution were required to produce a permanent lather?

How does washing soda act in softening water of temporary hardness?

6. To the remaining test-tube add 5 ml. of limewater and observe any changes. Test with soap solution as in Procedure 3.

Describe and account for any changes observed.

Write the equation for the reaction.

How many drops of soap solution were required to produce a permanent lather? How does this number compare with the number of drops required in each of the above Procedures?

Does the addition of limewater soften water of temporary hardness as efficiently as (a) boiling, and (b) the addition of washing soda?

QUESTIONS

1. Write equations to show how water of temporary hardness may be prepared in the laboratory.

2. What is meant by "water of temporary hardness"?

3. From the above experiments name four ways by which temporary hardness may be removed from water. Which method is the most effective? Why?

EXPERIMENT 101. Water of permanent hardness.

1. Quarter fill a test-tube with distilled water and add one drop of soap solution. Cover the mouth of the tube with your thumb and

shake the tube vigorously back and forth ten times. Observe the effect of the soap on the distilled water. Continue to add soap solution, one drop at a time, shaking after each addition and counting the number of drops, until a permanent lather is produced.

What effect is produced when soap solution is added to distilled water?

How many drops of soap solution were required to produce a permanent lather?

2. Quarter fill two test-tubes with distilled water. To one test-tube add a piece of calcium chloride the size of the head of a common pin, and to the other test-tube add a similar sized piece of magnesium sulphate. Add one drop of soap solution to each test-tube and shake each test-tube vigorously ten times. Continue to add soap solution, one drop at a time, shaking each time and counting the number of drops, until a permanent lather is produced. Observe all the changes that take place.

Describe and account for the change that takes place when soap is added to a distilled water solution of (a) calcium chloride, and (b) magnesium sulphate.

How many drops of soap solution were required to produce a permanent lather in each case?

In what respect is permanent hardness of water similar to temporary hardness?

What general reaction takes place whenever water is softened?

Does the addition of soap soften permanently hard water? Explain.

3. Quarter fill a test-tube with distilled water and add a piece of calcium chloride the size of the head of a common pin. Heat the solution to the boiling point, allow the test-tube to cool, and test the solution with soap solution as in Procedure 2.

How many drops of soap solution were required to produce a permanent lather?

Does boiling water containing calcium chloride produce any visible change?

What effect has boiling water of permanent hardness on the hardness of the water?

4. Quarter fill a test-tube with distilled water and add a piece of magnesium chloride the size of a head of a common pin. To this solution add a crystal of washing soda the size of a bean, shake the test-tube, and observe any changes that take place. Test the resulting mixture with soap solution to determine the number of drops required to produce a permanent lather.

What visible change takes place when a washing soda crystal is added to the magnesium chloride solution?

Account for the action of the sodium carbonate and write the equation to express the reaction that takes place.

How many drops of soap solution were required to produce a permanent lather?

Why is washing soda an effective water softener?

QUESTIONS

1. List all the substances whose solutions cause water to be hard.
2. What is a common method of detecting hardness in water?
3. Discuss the relative merits of the terms: "temporary" or "carbonate" hardness; "permanent" or "non-carbonate" hardness.
4. What advantage is gained by using chemical water softeners when softening water?

Appendix

A. IMPORTANT UNITS IN THE METRIC SYSTEM

Length

10 millimetres (mm.)	= 1 centimetre (cm.)
10 centimetres	= 1 decimetre (dm.)
10 decimetres	= 1 metre (m.)
1000 metres	= 1 kilometre (km.)

Mass (Weight)

10 milligrams (mg.)	= 1 centigram (cg.)
10 centigrams	= 1 decigram (dg.)
10 decigrams	= 1 gram (gm.)
1000 grams	= 1 kilogram (kg.)

Volume

1000 cubic centimetres (cc.)	= 1 litre (l.)
1 cc.	= 0.001 litres = 1 millilitre (ml.)

While a cubic centimetre and a millilitre have not exactly the same value, the variation is so slight that for practical purposes these two units may be considered identical.

B. RELATIONSHIP OF SOME ENGLISH AND METRIC UNITS

1 inch (in.)	= 2.54 centimetres
1 metre	= 39.37 inches
1 kilometre	= 0.62 miles
1 pound (lb.) Avoirdupois	= 453.6 grams
1 ounce (oz.) Avoirdupois	= 28.35 grams
1 kilogram	= 2.20 pounds
1 litre	= 0.88 quarts (British)
1 cubic foot (cu. ft.)	= 28.32 litres

C. TABLE OF SOLUBILITIES

(In cold water)

S	= soluble in water
M	= slightly or moderately soluble in water
I	= insoluble or nearly insoluble in water

— = nonexistent compound, or decomposed by water

R = reacts chemically with water forming a base

	Acetate	Bromide	Carbonate	Chlorate	Chloride	Hydroxide	Iodide	Nitrate	Oxide	Phosphate	Sulphate	Sulphide	Sulphite
Aluminum.....	S	S	—	S	S	I	S	S	I	I	S	—	—
Ammonium.....	S	S	S	S	S	S	S	S	—	S	S	S	S
Barium.....	S	S	I	S	S	M	S	S	R	I	I	S	M
Calcium.....	S	S	I	S	S	M	S	S	R	I	M	M	M
Copper (—ic)...	S	S	I	S	S	I	S	S	I	I	S	I	—
Iron (—ous)...	S	S	I	—	S	I	S	S	I	I	S	I	M
Iron (—ic).....	S	S	—	—	S	I	—	S	I	I	S	I	—
Lead.....	S	M	I	S	I	I	I	S	I	I	I	I	I
Magnesium....	S	S	I	S	S	M	S	S	R	I	S	I	M
Manganese....	S	S	I	—	S	I	—	S	I	I	S	I	—
Mercury(—ous)	S	I	I	S	I	—	I	S	I	I	M	—	—
Mercury (—ic)...	S	M	I	S	S	—	I	S	I	I	S	I	—
Potassium.....	S	S	S	S	S	S	S	S	R	S	S	S	S
Silver.....	M	I	I	S	I	—	I	S	I	I	M	I	S
Sodium.....	S	S	S	S	S	S	S	S	R	S	S	S	S
Strontium.....	S	S	I	S	S	M	S	S	R	I	I	M	—
Zinc.....	S	S	I	S	S	I	S	S	I	I	S	I	I

D. GENERAL RULES FOR SOLUBILITY

Certain generalizations can be made concerning the compounds listed in the accompanying table. Exempting a few unimportant exceptions, the following statements hold true:

1. All *potassium*, *sodium*, and *ammonium* compounds are soluble in water.
2. All *acetates*, *chlorates*, and *nitrates* are soluble.
3. All *chlorides*, *bromides* and *iodides* are soluble, except those of *lead*, *mercury* (*mercurous*), and *silver*. (Lead chloride is noticeably soluble in *hot* water.)
4. All *sulphates* are soluble, except those of *barium*, *lead*, and *strontium* (The sulphates of *calcium* and *silver* are only slightly soluble.)
5. All normal *carbonates* and *phosphates* are insoluble, with the exceptions of the *potassium*, *sodium* and *ammonium* compounds.
6. All *oxides* are insoluble, except those of *potassium*, *sodium*, *barium*, *calcium*, *magnesium*, and *strontium*. These oxides react with water producing hydroxides which dissolve in water.
7. All *sulphides* are insoluble, except those of *potassium*, *sodium*, *ammonium*, *barium*, *strontium*, and *calcium*. (Calcium sulphide and strontium sulphide are only moderately soluble.)
8. All *hydroxides* are insoluble, except those of *potassium*, *sodium*, *ammonium*, *barium*, *calcium*, *magnesium*, and *strontium*. (The hydroxides of *barium*, *calcium*, *magnesium* and *strontium* are only slightly soluble.)

E. ACTIVITY TABLE

(Some of the common metals)

1. Potassium	9. Tin
2. Sodium	10. Lead
3. Barium	11. Hydrogen
4. Calcium	12. Copper
5. Magnesium	13. Mercury
6. Aluminum	14. Silver
7. Zinc	15. Gold
8. Iron	16. Platinum

The relative activity with which metals displace *hydrogen* from compounds such as water and dilute acids is shown by this table. Metals *below* hydrogen will not displace hydrogen from these compounds.

F. GENERAL PROPERTIES OF METALS AND NON-METALS*Metals***PHYSICAL PROPERTIES**

1. Good conductors of heat.
2. Good conductors of electricity.
3. Greyish metallic lustre, except copper (red) and gold (yellow).
4. Malleable. (Gold may be pounded into sheets $\cdot 0001$ cm. thick.)
5. Ductile. (Wire may be produced weighing $\cdot 000005$ gm/cm.)

CHEMICAL PROPERTIES

1. All form oxides either by direct combination or by some intermediate reaction.
2. Metallic oxides are basic anhydrides.
3. The more active metals (those above Hydrogen in the activity series) react with acids liberating hydrogen.
4. The more active metals react with water forming hydrogen and a base.
5. Metals react with non-metals to form salts.

*Non-Metals***PHYSICAL PROPERTIES**

1. Poor conductors of heat.
2. Poor conductors of electricity, except graphite and selenium.
3. Vary considerably in colour; sulphur (yellow), phosphorus (white or yellow or red), chlorine (greenish-yellow), bromine (brownish red), iodine (steel-grey with violet vapour).
4. Not malleable.
5. Not ductile.

CHEMICAL PROPERTIES

1. All form oxides either by direct combination or by some intermediate reaction.
2. Non-metallic oxides are acid anhydrides.
3. Only the most active non-metals react with water and when they do, they form acids and liberate oxygen.
4. Non-metals react with metals to form salts.

G. PRESSURE OF WATER VAPOUR, OR AQUEOUS TENSION

(In millimetres of mercury)

TEMPER- ATURE	PRESSURE	TEMPER- ATURE	PRESSURE	TEMPER- ATURE	PRESSURE
0°C.	4.6 mm.	11°C.	9.8 mm.	22°C.	19.8 mm.
1°	4.9 mm.	12°	10.5 mm.	23°	21.0 mm.
2°	5.3 mm.	13°	11.2 mm.	24°	22.3 mm.
3°	5.7 mm.	14°	12.0 mm.	25°	23.7 mm.
4°	6.1 mm.	15°	12.8 mm.	26°	25.1 mm.
5°	6.5 mm.	16°	13.6 mm.	27°	26.7 mm.
6°	7.0 mm.	17°	14.5 mm.	28°	28.3 mm.
7°	7.5 mm.	18°	15.5 mm.	29°	29.9 mm.
8°	8.0 mm.	19°	16.5 mm.	30°	31.7 mm.
9°	8.6 mm.	20°	17.5 mm.	50°	92.3 mm.
10°	9.2 mm.	21°	18.6 mm.	100°	760.0 mm.

When a gas is collected over water the true pressure exerted by the gas itself is obtained by subtracting from the observed barometric pressure the pressure due to water vapour at the temperature of measurement.

H. DENSITY OF GASES

(At 0°C., and 760 mm. pressure)

Based on *approximate* atomic weights

Name	Form- ula	Mol. Wt.	Density (gm./litre)
Acetylene	C ₂ H ₂	26	1.16
Air	1.29
Ammonia	NH ₃	17	0.76
Argon	A	39.94	1.78
Carbon dioxide	CO ₂	44	1.96
Carbon monoxide	CO	28	1.25
Chlorine	Cl ₂	71	3.17
Fluorine	F ₂	38	1.70
Helium	He	4	0.18
Hydrogen	H ₂	2	0.09
Hydrogen bromide	HBr	81	3.61
Hydrogen chloride	HCl	36.5	1.63

Name	Form- ula	Mol. Wt.	Density (gm./litre)
Hydrogen iodide	HI	128	5.71
Hydrogen sulphide	H ₂ S	34	1.52
Krypton	Kr	83.7	3.73
Methane	CH ₄	16	0.71
Neon	Ne	20	0.90
Nitric oxide	NO	30	1.34
Nitrogen	N ₂	28	1.25
Nitrogen dioxide	NO ₂	46	2.05
Nitrous oxide	N ₂ O	44	1.96
Oxygen	O ₂	32	1.43
Sulphur dioxide	SO ₂	64	2.86
Xenon	Xe	131	5.85

I. APPROXIMATE ATOMIC WEIGHTS OF THE COMMON ELEMENTS

NAME	SYMBOL	ATOMIC WEIGHT	NAME	SYMBOL	ATOMIC WEIGHT
Aluminum	Al	27	Lead	Pb	207
Antimony	Sb	122	Magnesium	Mg	24
Arsenic	As	75	Manganese	Mn	55
Barium	Ba	137	Mercury	Hg	200
Bromine	Br	80	Nitrogen	N	14
Calcium	Ca	40	Oxygen	O	16
Carbon	C	12	Phosphorus	P	31
Chlorine	Cl	35.5	Potassium	K	39
Copper	Cu	63.6	Silver	Ag	108
Fluorine	F	19	Sodium	Na	23
Gold	Au	197	Sulphur	S	32
Hydrogen	H	1	Tin	Sn	119
Iodine	I	127	Zinc	Zn	65
Iron	Fe	56			

J. PREPARATION OF LABORATORY REAGENTS

Acetic Acid (dilute)	: 1 vol. 'glacial with 2½ vols. water.
Hydrochloric acid (conc.)	: use the C.P. acid of commerce.
Hydrochloric acid (dilute)	: equal vols. of C.P. acid and water.
Nitric acid (conc.)	: use the C.P. acid of commerce.
Nitric acid (dilute)	: 1 vol. C.P. acid with 2 vols. water.

Sulphuric acid (conc.)	: use the C.P. acid of commerce.
Sulphuric acid (dilute)	: 1 vol. C.P. acid with 5 vols. water.
Ammonium hydroxide (conc.)	: use the C.P. product of commerce.
Ammonium hydroxide (dilute)	: 1 vol. C.P. product with $1\frac{1}{2}$ vols. water.
Potassium hydroxide (conc.)	: use a saturated solution.
Potassium hydroxide (dilute)	: dissolve 336 gm. of solid potassium hydroxide in water, and dilute to 1 litre.
Sodium hydroxide (conc.)	: use a saturated solution.
Sodium hydroxide (dilute)	: dissolve 240 gm. of solid sodium hydroxide in water, and dilute to 1 litre.
Silver nitrate solution	: dissolve 34 gm. of solid silver nitrate in water and dilute to 1 litre.

K. SULPHURIC ACID—DICHROMATE CLEANING SOLUTION

Cautiously (TECH. 3) pour 200 ml. of concentrated sulphuric acid into 150 ml. of cold water stirring continuously, and without further heating saturate the hot solution with powdered potassium dichromate.

To clean glass vessels, fill them with this cold solution and allow to stand overnight or longer.

L. SOME COMMON NAMES OF SUBSTANCES AND THEIR FORMULAE

COMMON NAME	FORMULA
Alum	$K_2SO_4 \cdot Al_2(SO_4)_3 \cdot 24H_2O$
Ammonia Water	NH_4OH
Aqua regia	HCl (4 volumes) + HNO_3 (1 volume)
Baking powder	$NaHCO_3$ + starch + an acid-acting ingredient
Baking soda	$NaHCO_3$
Bleaching powder	$CaOCl_2$
Blue vitriol (bluestone)	$CuSO_4 \cdot 5H_2O$
Bone black	C
Caustic potash	KOH
Caustic soda	NaOH
Chile saltpetre	$NaNO_3$
Chloride of lime	$CaOCl_2$
Coke	C
Common salt	NaCl

Cream of tartar	$\text{KHC}_4\text{H}_4\text{O}_6$
Dextrose (glucose)	$\text{C}_6\text{H}_{12}\text{O}_6$
Diamond	C
Dry ice	CO_2 (solid)
Epsom salts	$\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$
Flowers of sulphur	S
Glauber's salt	$\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$
Glucose (dextrose)	$\text{C}_6\text{H}_{12}\text{O}_6$
Glycerine (glycerol)	$\text{C}_3\text{H}_5(\text{OH})_3$
Grain alcohol (ethyl)	$\text{C}_2\text{H}_5\text{OH}$
Graphite	C
Gypsum	$\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$
Household ammonia	NH_4OH
Hypo (crystals)	$\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$
Lampblack	C
Laughing gas	N_2O
Lime (hydrated)	$\text{Ca}(\text{OH})_2$
Lime (quicklime)	CaO
Lime (slaked)	$\text{Ca}(\text{OH})_2$
Limestone	CaCO_3
Limewater	$\text{Ca}(\text{OH})_2$
Lye	NaOH
Marble	CaCO_3
Marsh gas (methane)	CH_4
Methyl alcohol (methanol)	CH_3OH
Muriatic acid (hydrochloric)	HCl
Napthalene (moth balls)	C_{10}H_8
Oil of vitriol	H_2SO_4
Plaster of Paris	$(\text{CaSO}_4)_2 \cdot \text{H}_2\text{O}$
Quicklime	CaO
Quicksilver (mercury)	Hg
Sal ammoniac (ammonium chloride)	NH_4Cl
Sand	SiO_2
Slaked lime	$\text{Ca}(\text{OH})_2$
Soap (sodium stearate)	$\text{C}_{17}\text{H}_{35}\text{COONa}$
Soda ash	Na_2CO_3
Sugar (sucrose)	$\text{C}_{12}\text{H}_{22}\text{O}_{11}$
Vinegar (acetic acid)	$\text{HC}_2\text{H}_3\text{O}_2$
Washing soda (sal soda)	$\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$
Wood alcohol	CH_3OH

M. FIRST AID IN THE CHEMISTRY LABORATORY

Faulty technique is one of the chief causes of accidents and, because it involves the human element, is one of the most difficult with which to cope. Good technique demands that students maintain an awareness of the danger in handling chemicals, are able to assemble apparatus correctly, handle materials properly, and proceed methodically. Proper methods of mixing acids, cutting glass tubing and inserting it through stoppers, etc., are as important as correct experimental results. In spite of proper precautions, accidents do occur from time to time and the fundamental rule to remember is to *KEEP CALM*. Make certain that the right person begins first aid, and in the case of serious accidents have *one* person call the school nurse or a physician immediately.

A supply of first-aid materials, kept in an easily accessible cabinet or cupboard, should be present as standard equipment in every laboratory to treat the accidents which sometimes occur. The following items are of value:

FIRST-AID MATERIALS

Rolls of 1, 2, and 3-inch sterile gauze bandage

Sterile gauze, 3-inch squares

Absorbent cotton, sterilized

Adhesive tape—to keep bandages in place

Band-aids

Eye cup

Scissors, tweezers, wood applicators, medicine droppers

Tincture of iodine (3.5% solution)

Petroleum jelly (vaseline)

Sodium bicarbonate solution (5%)

Boric acid solution (20 gm. in 1 litre of water)

Boric acid ointment

Butesin picrate ointment

Carron oil (equal volumes of limewater and linseed or cottonseed oil)

Castor oil or olive oil

Aromatic spirits of ammonia

Tincture of merthiolate

Naphtha gasoline or grain alcohol

Fire blanket

BURNS

Acid in the eyes: Pull back the lids and flood the eye thoroughly with running water. Then bathe the eye with a 5% sodium bicarbonate solution, using an eye cup. Dry with sterile gauze and treat

the eye with a few drops of olive oil or castor oil which acts as a soothing agent.

Base (alkali) in the eyes: Flood the eye thoroughly with running water. Then bathe the eye with a 5% boric acid solution, using an eye cup. Dry with sterile gauze and add a few drops of olive oil or castor oil.

Acid burns on the skin: Wash the burned area thoroughly with running water and then with a 5% solution of sodium bicarbonate. Apply sterile petroleum jelly or carron oil or butesin picrate ointment, and protect with gauze bandage.

Alkali burns on the skin: Wash thoroughly with running water and neutralize with a boric acid solution. Wash again and apply boric acid ointment or sterile petroleum jelly. Bandage the burned area loosely.

Heat burns: Cover the burned area with butesin picrate ointment or petroleum jelly and then bandage loosely with sterile gauze. *Do not break any blister formed.* Do not use tincture of iodine on *any burn* because of its tendency to increase skin damage.

Bromine Burns: Wash thoroughly with running water and immediately treat with a 5% solution of sodium bicarbonate. Dry and apply carron oil or keep the burned area moist by applying glycerine.

Phosphorus Burns: Flood with water. Then apply a 1% aqueous solution of copper sulphate. If no phosphorus remains in the burn, treat as for "heat burns" described above.

INHALING TOXIC FUMES

Remove the victim to outside air and encourage him to take deep breaths of fresh air. *Acid vapours* (chlorine, bromine, hydrogen chloride, etc.) may be treated by having the patient inhale the vapour from dilute ammonium hydroxide (aromatic spirits of ammonia). If the gas is *ammonia*, allow him to inhale the vapour from a dilute solution of acetic acid.

Headaches or dizziness are nature's warnings of toxic atmospheres. Fresh air, five grains of aspirin, and a half hour of rest will usually relieve toxic headaches.

CUTS

If bleeding is copious it must first be stopped. An arterial cut is indicated by scarlet blood flowing intermittently. Use a piece of rubber tubing ($\frac{1}{2}$ -inch bore) as a tourniquet between the heart and the cut. Never tighten more than necessary to stop the bleeding and do not maintain the pressure for longer than fifteen minutes.

at a time. Loosen it and allow blood to flow for a few seconds and then retighten. Loosen, but do not remove the tourniquet as the blood clots. *Send student to a physician.*

The blood which pours from veins is dark red in colour and appears in a steady flow from the wound, without the pulsations or "spurts" seen in an arterial cut. Venous bleeding is more readily controlled. Apply a sterile gauze compress directly over the wound and apply firm pressure with the fingers until the bleeding stops or a clot forms. When the bleeding ceases, do not disturb the compress, and bandage with sterile gauze.

Minor cuts are serious, because of possible ingress of bacteria leading to infection. Wash the wound zone with a gauze compress saturated with naphtha gasoline or grain alcohol to remove any grease. If foreign particles are present on the surface, they should be removed by gentle brushing with sterile gauze. Apply an antiseptic solution (tincture of iodine, tincture of merthiolate, mercurochrome solution, etc.), and bandage with sterile gauze.

Patients with deep puncture wounds, as from glass tubing, broken thermometers, etc., should be sent to a physician.

ACCIDENTS TO CLOTHING

Acids on clothing. Apply limewater or a dilute solution of sodium bicarbonate to the affected area to neutralize the acid. Then wash well in running water.

Bases on clothing. Neutralize the alkali by applying dilute acetic acid to the affected area. Then wash thoroughly with water.

Clothing on fire. Smother the fire (exclude oxygen) with a woollen blanket. A suit coat may be used if the laboratory is not equipped with a fire blanket. Speed is essential if severe injury is to be avoided.

FAINTING

Place the patient in a comfortable lying position with the head lower than the rest of the body. Open the windows and insist that onlookers do not crowd around the patient, since a good circulation of air is imperative. The return to consciousness can be accelerated by having the patient breathe the vapour from aromatic spirits of ammonia placed on a folded handkerchief. Placing a towel moistened with cold water on the patient's forehead is also helpful. When consciousness is restored, the patient should be kept lying down until fully recovered and should then be allowed to leave the classroom only when accompanied by reliable assistants.

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